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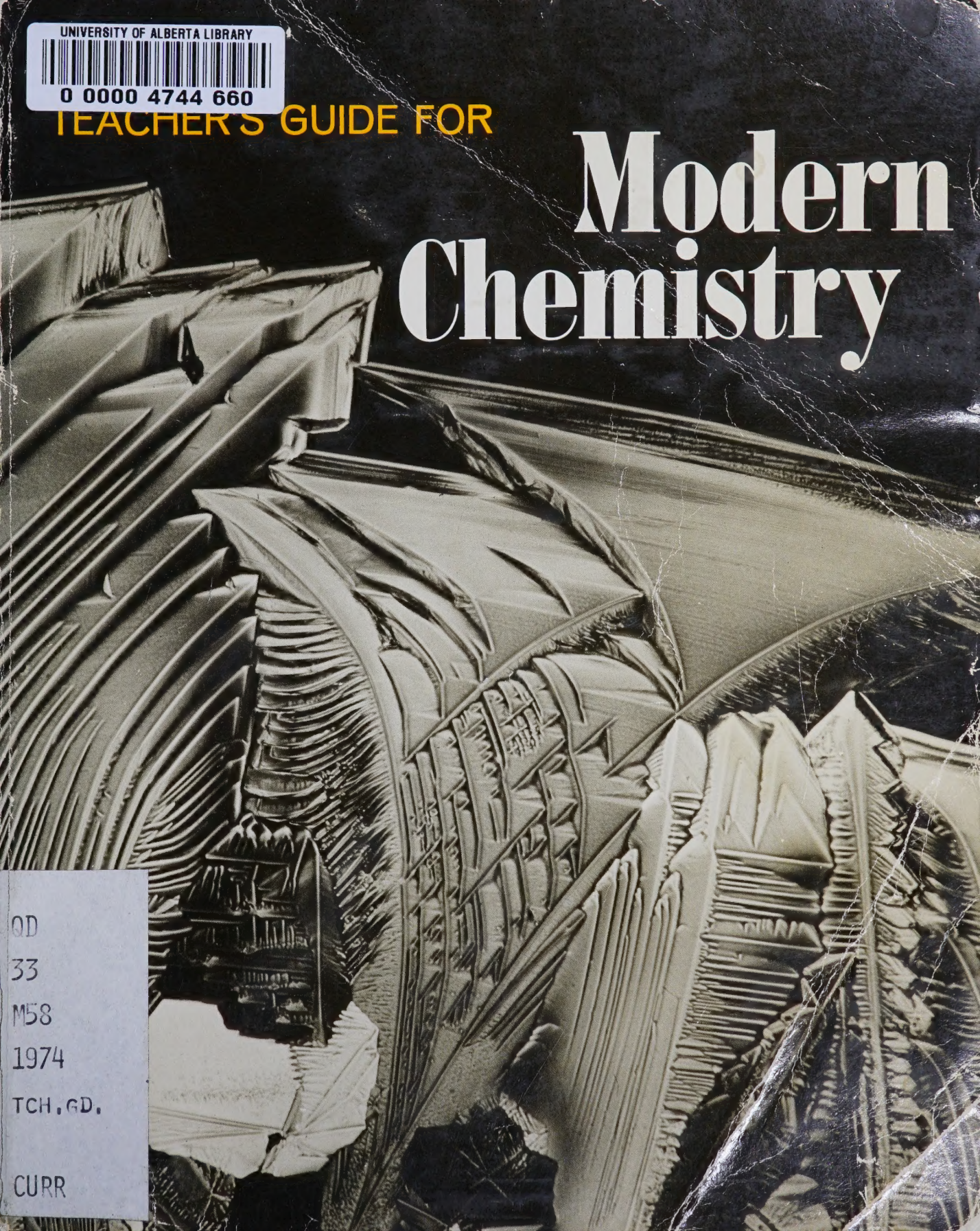
TEACHER'S GUIDE FOR

Modern Chemistry



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Teacher's Guide to

Modern Chemistry

H. Clark Metcalfe
John E. Williams
Joseph F. Castka

HOLT, RINEHART AND WINSTON, INC.
New York • London • Toronto • Sydney

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Introduction

The *Modern Chemistry* program is intended to help you, the teacher in your effort to achieve the objectives of science teaching. The program consists of a variety of components which allow the teacher ample approaches for an interesting and challenging course.

The Holt *Modern Chemistry* Program
Metcalfe, Williams, and Castka
Modern Chemistry (student text)
Teacher's Guide to Modern Chemistry
Laboratory Experiments in Chemistry
Exercises and Experiments in Chemistry
Teacher's Edition to Exercises and Experiments in Chemistry
Tests in Chemistry
Alternate Tests in Chemistry

This teacher's guide offers you aid in using the complete program. The *Preface* of *Modern Chemistry* 1974 ennumerates the changes from previous editions. The experiments mentioned in this teacher's guide, unless otherwise identified, are from *Laboratory Experiments in Chemistry.* These experiments are identical with those in *Exercises and Experiments in Chemistry* with the exception of write-in spaces.

The experiments by categories are:

1. Descriptive-observational, in which students observe, summarize, and draw conclusions.
2. Qualitative, in which students record observations in data tables and draw conclusions on the bases of qualitative comparison.
3. Quantitative, in which students measure various quantities and make computations based on these quantitative observations.

Most of the other recommended experiments are from *Scientific Experiments in Chemistry*. These are of the "open-ended" type and are identified by name and number.

The exercises in *Exercises and Experiments in Chemistry* may be used in a variety of ways as instructional aids. They may be assigned for completion after the student has read the chapter, or assigned as a self-test or as a review before a test. The items in each exercise parallel the text in sequence so that the conscientious student can easily locate the answer in the text. The exercises also serve as instructional complements to the questions and problems in the text.

The *Teacher's Edition to Exercises and Experiments in Chemistry* includes answers and other useful details for the laboratory section of the program.

Tests in Chemistry and *Alternate Tests in Chemistry* parallel each other. They provide different but equivalent tests for classes meeting at different times during the day. Keys are provided for complete tests.

The 1974 version of the *Teacher's Guide to Modern Chemistry* consists of Part I, An Overview of the Teaching of Chemistry; Part II, Teaching Suggestions for a Modified Chemistry Course; Part III, Answers to Questions and Problems of the Student Text; Part IV, Bibliography; Part V, Key to Suppliers; Part VI, Word Lists; Part VII, Teacher's Notes.

The Teacher Note section is marked with the chapter section of the student text to which they pertain. Included in the notes are: appropriate demonstrations, audiovisual aids, and classroom activities. The audiovisual aids are organized as films, loops, filmstrips, and transparencies. The suppliers are indicated by the code letters to the left of the entry. (The supplier's name can be found in Part V.) The content of some of the aids is apparent from the placement within the Teacher Note. The running time of the aid and type is provided for example (11 min., 16mm, 8m, s-8).

The Bibliography consists of an extensive listing of reference books. A chart of references is also provided. It correlates pertinent sections of the reference with a given chapter of *Modern Chemistry*, 1974.

Part I

An Overview of the Teaching of Chemistry

The International Conference on Education in Chemistry held at Snowmass-at-Aspen, Colorado, during July 1970, was sponsored by the Division of Chemical Education of the American Chemical Society in collaboration with the Committee on the Teaching of Chemistry of the International Union of Pure and Applied Chemistry.[1] The conference considered the problems and possibilities for chemical education in the 1970's. During 1970, consequent meetings were held at Chicago in September and at Colorado State University in December.[2]

The following quotations from the Snowmass conference indicate the problems and suggest directions that correctional processes must take.[3] "Somehow, in the eyes of the layman, science and scientists, as related to the pressing ills of contemporary society, have become part of the problem, rather than leaders in the solution."

"For some reason, the public has not caught the spirit—the essential humaneness—of the scientific enterprise. Science is regarded as a coldly intellectual or crassly technological ogre, not as a warmly human endeavor, geared to the serving of human needs."

"This failure can be laid chiefly at the doorstep of scientists themselves, for their failure to communicate adequately—to the young (in the educational process) and to the adult lay public, in the use of the mass communications media."

The Colorado State Conference has established plans and criteria for a new, realistic program for the general student in the high schools (and the two year colleges). These include suggestions such as: Meet the student at the level of his own experience, interest, concerns, and problems. Make the curriculum flexible. Develop the program so that it consists of integrated modular units of two types—motivational and conceptual. Base the motivational modules on topics of real and current interest to students—energy sources, body chemistry, population control, air and water pollution, psychedelic drugs. These motivational modules can be supported by conceptual modules. The conceptual modules develop underlying scientific principles basic to the motivational modules. These principles include kinetic molecular theory, mole concept, atomic and molecular structure, the periodic table, and oxidation and reduction. Coordinate a particular motivational module with one or more basic modules. The learning of basic concepts is reinforced by considerable overlapping. The needs and interests of the students would determine the selection of the modules.

There is every reason to believe that the experimentation, underway during the early 1970's, will continue in attempts to meet the problems posed and guidelines suggested by meetings such as those held at Snowmass and subsequent conferences. Experimental and emergent curricula such as the one indicated by the title of an article in *The Science Teacher*, "Individualizing Chemistry—A Method Used in an Open High School,"[4] reveal attempts to deal with current educational trends. "Environmental Chemistry" is the

[1] "Preliminary Report–International Conference on Education in Chemistry," Snowmass-at-Aspen, *Journal of Chemical Education,* 48, 3 (1971).

[2] Vander Werf, C. A., Journal of Chemical Education, 48, 91 (1971); Young, D. P., *Journal of Chemical Education,* 48, 725 (1971).

[3] Vander Werf, C. A., *Journal of Chemical Education,* 48, 91 (1971).

[4] Kline, A. A., TST, 39, 61 (1972).

If chemistry were just distillation, explosives and boy-girl relationships, I'd have the most popular course in the school.

theme of a series of nine articles in the Journal of Chemical Education.[5] The "Report of the Curriculum Committee of the Division of Chemical Education, American Chemical Society,"[6] states that some 200 members of the committee are attempting to develop outlines and statements of behavioral objectives and to integrate them for use at the high school and college levels. This "position" paper discusses the need to convey to the student what role science plays in our culture and in his own life. It also makes the statement that although students rarely remember more than a few facts, "All students leave our courses with long-remembered attitudes." It also warns that the emphasis on behavioral objectives, legitimate as it is, should not be regarded as the new "ultimate solution to all problems."

Curriculum experimentation also continues in the area of integrated physics and chemistry courses. Interest in the unification of the physical and biological sciences is indicated by the titles of such articles as: "The Unified Science Study Program."[7] and "Interdisciplinary Science Education-Report of a Conference."[8]

To summarize, curriculum developments during the 1970's focus on:

1. Relevance: Scientific Literacy, "Humanizing" chemistry.
2. Flexibility.
3. Individualizing instruction.
4. Interdisciplinary science.

TEACHING THE COURSE

Modern Chemistry 1974 provides the basic facts and concepts reviewed in a framework of laboratory and textual content. They are aimed at developing the related understandings in greater depth. In a similar vein, equation writing is associated with the mole concept and the related mathematics. Formula writing involves understanding of bonding, the application of oxidation-number rules, and of total positive and negative charge agreement.

A large number and variety of "newer" tools are presently available for strategic use in teaching chemistry. The suggestions for their use are specifically detailed in teacher notes.

Starting with 1970, the September issues of the *Journal of Chemical Education* have a section (in the Book Buyers' Guide) devoted exclusively to audio-visual materials including cartridged filmloops, filmstrips, slides, and transparencies. Articles which contain suggestions for chemistry teachers interested in individualization and "open schools," although they are not specifically secondary school nor chemistry-oriented, appear quite frequently. Such a series appeared in the December 1970 issue of *The Science Teacher.*

Teaching the course therefore requires a realignment of many of the "old" demonstrations, techniques, and resources, synthesized with the "new." Able students should be encouraged to attend Summer Institutes, participate in Science Fairs, and take part in such competitions as the Annual Science Talent Search and

[5] "Environmental Chemistry," *Journal of Chemical Education,* 49, 5–30 (1972).

[6] Cassidy, H. J., *Journal of Chemical Education,* 49, 34 (1972).

[7] Morgan, S., TST, 37, 41 (1970).

[8] Supplement, TST, 37, No. 8 (1970) Scubert L., Conference Director.

the Annual NASA Science Congress. However, we must not lose, nor lose sight of, the chemistry student who may be a commercial student, a general student, or one whose academic career terminates at the high school level. The *Modern Chemistry* program provides sufficient diversity to meet the needs of such students. The descriptive chemistry and chemical technology appear largely in the latter half of the textbook. Judicious selection from among the more theoretical-mathematical first half provides the student with the fundamentals needed to cope with important concepts, principles, and processes, and to understand the role of chemistry in the world of today. See Part II of this Teacher's Guide.

Lesson Planning

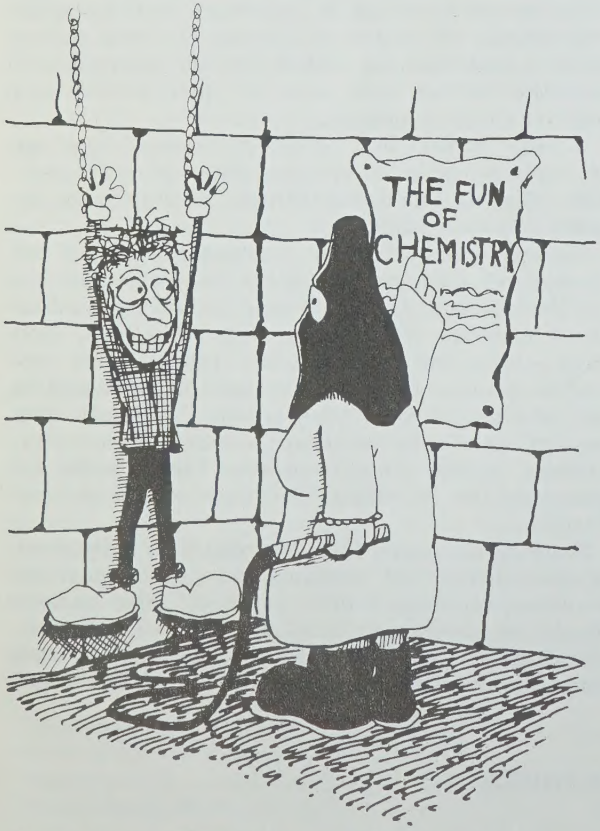
The teacher is not a taskmaster, technician, transmitter of information, nor trainer for passing examinations. In the finest sense he is a creative artist, an engineer of ideas, an architect of character. The teacher's primary tool is careful planning, over-all and day-by-day. Lesson plans generally mean plans for a single day or period. They are parts of topic plans or unit plans. In making the daily lesson plan, the teacher asks himself Why? What? and How? The lesson plan starts with the *why*, the objectives of science education which depend on educational philosophy. The *what* deals with the learnings (the topics) which are the more immediate aim of a particular lesson. The *how* is the method.

In lesson planning, the determination of sequence and timing play an important part. The sequence of induction should proceed from demonstration to observations to inferences, explanations, generalizations, and ultimate solution of the problem. Student participation should be encouraged. The more deeply students are involved in planning, the greater is their drive toward learning. However, the teacher must always realize that students have limited experience. They require motivation so that the lesson aim becomes the student aim.

Lesson Methods

APPROACH AND MOTIVATION

Most textbooks consist of facts, generalizations and concepts, organized in logical and systematic fashion.



Say what you will about her methods, Ms. Snidely has a very low failure rate.

Mr. Walanda goes to great lengths to hold the students' attention.

A lesson based mainly on the textbook exposition frequently arouses little student interest. It brings a lukewarm response, and has questionable success in promoting learning and retention.

The skilled teacher carefully selects the way in which units or subunits are introduced and developed. To translate the textbook into an effective learning experience, the teacher must challenge the student and make him fully aware of the value of each lesson. The emotional drives which initiate and maintain human endeavor must be utilized to the utmost.

Since interest is a highly personal and individual matter, not all motivations will reach all students. The teacher must decide which procedures will probably reach the maximum number of individuals.

By and large all motivating devices present a problem, or lead to the recognition of problems. The approach and motivation should not be merely a "gimmick" to start the ball rolling. At some time during the lesson, the motivation must be reconsidered. Motivation may have to be reinforced during certain lessons. Finally, continued use of one type of motivation palls on the student and the device becomes mere routine instead of inspiration. Try for variety in approach and motivation as well as in general method.

QUESTIONING

Questioning is the usual process by which the effectiveness of teaching is determined. Good technique is important in both the framing and the administration of the questions. Timeliness, timing, sequential placement, and method must all be considered.

Portia Mason really knows how to question students.

Skillful questioning does not cease with a correct answer nor aim merely at the revision of an incorrect one. Further questioning which reveals the thinking that led to the response will help the student revise an initial incorrect answer. Explanation of how a correct answer was arrived at has instructive value for the rest of the class. The continued questioning may also result in greater depth and scope of understanding for all.

THE ASSIGNMENT

Generally, an assignment consists of specific reading in a single text and related questions, the answers to which constitute the written assignment. Such questions may be selected from those supplied with the textbook, or may be framed by the teacher.

This formal type of assignment has disciplinary value in that students are responsible for achievement and effort. It has many variations which improve its value. Some teachers supplement page assignments by writing down the topic or problem, sometimes indicating important subtopics. However the assignment is given, some oral comment on it is required.

Training in techniques of library research such as note-taking, abstracts, etc. is helpful. However, this procedure should seldom be used with lessons calling for the development of fundamental skills, such as equation writing or mathematical problems. Many less able students cannot make the necessary connections between textbook explanations or examples and a task based upon such materials. Written assignments in such areas should be based on concepts and skills developed in the classroom.

The ideal assignment is one chosen by the student himself. It may arise from teaching situations in which no formal reading or written tasks have been given. Students may inquire where or how they may obtain information about a particular aspect of a problem. The class may decide to list the appropriate problems, information, demonstration, charts, apparatus, and projects needed for further study. Each student or specific committee chooses the particular assignment in which interest is greatest. This process results in self-sustained motivation. The teacher suggests, guides, directs, and constructively criticizes.

Such assignments have additional values. Students are encouraged to consult a variety of textual and reference materials. Teacher planning may involve the assembling of a bibliography. Students may be asked to add to this bibliography to expand resources for future groups. The teacher may choose to pinpoint various sections or chapters of specific references and organize such references into subtopical units. As their abilities develop, students are given greater responsibility for the location and choice of reference materials. Skill in

using the index and table of contents, outlining, and summarizing is a valuable outgrowth of this method of assignment. In general, student and group identification with the assignment makes it more meaningful.

Lesson Types

Methods will vary in terms of the specific outcomes desired, the specific school environment, the individual needs of the student, and the availability of instructional materials. The teacher uses methods which achieve the objectives he desires.

As a broad plan of organization, the unit-problem method has gained general acceptance. In brief, this method considers each unit as a broad problem. Analysis of the big unit problem leads to subsidiary problems, each of which is solved in turn, thereby leading to the ultimate solution of the unit problem. This is the main scheme upon which the teacher note portion is organized. In the following discussion of specific methods, applications to the unit-problem plan will be pointed out.

THE DEMONSTRATION METHOD

The demonstration-discussion or demonstration-developmental method, is more or less standard practice for the majority of classroom lessons in chemistry. These demonstrations are adapted to the logical development of content, the collection of information and facts, the development or verification of a generalization, principle, or concept, or the solution of a problem.

Mr. Booth makes chemistry seem so dramatic.

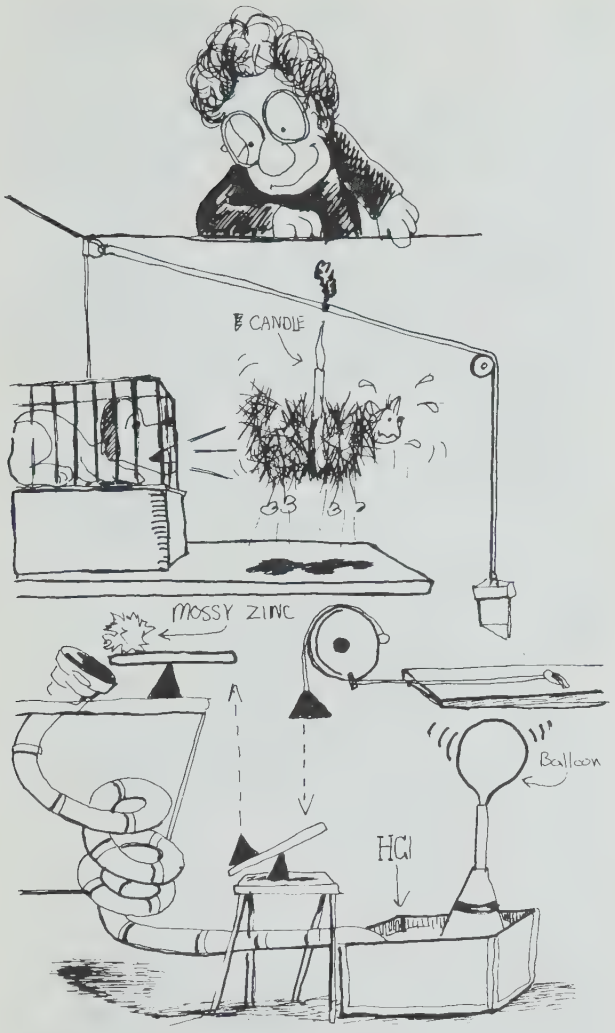
The ideal approach is the inductive one, in which the demonstration is preceded by a thorough discussion of the problem situation. Recognition of the problem should emerge from some motivating device or agency. Once the students recognize the problem, it should be clearly stated in simple language, preferably by a student. This procedure establishes the aim of the lesson and indicates the direction of further investigation. It may result in the emergence and evaluation of hypotheses and the tentative selection of an acceptable hypothesis. Properly directed, it will also lead to the suggestion of necessary controls, checks, and safety factors. Students should take part in as many phases of the demonstration as possible.

The objective of this method is not quantity of content coverage but rather quality of problem-solving skills and attitudes. The encouragement of the curious, questioning attitude is important. A good student question is much more important that a "correct" student answer. The chalk board may be used for the development or synthesis of diagrams representing the planned demonstration, and for recording procedural steps, important words, formulas, equations, observations, and conclusions as the demonstration proceeds. Large prepared charts, labeled or unlabeled, showing the apparatus, may be displayed at the appropriate time to supplement chalkboard diagrams.

This use of the demonstration-discussion method does not preclude the occasional use of other types of demonstrations. For example, the teacher or a prepared student group may present a series of significant demonstrations as an introduction to a unit, as a review, as a stimulating device, or as a training procedure. Occasioanlly, the "silent" type of demonstration may be performed. The demonstration is carried out in its entirety and the rest of the lesson is developed thereafter. A demonstration of this type may be performed with a substance like carbon dioxide. The preparation of the gas, method for its collection, and its properties are demonstrated. The students then list and/or discuss these properties and state the substantiating evidence. Since several of the properties of the gas, such as specific gravity, may be inferred from different aspects of the demonstration, this procedure permits students to gather and organize pertinent observations and data.

The following demonstration techniques are desirable and effective:

1. Check the apparatus, materials, and order of operations before the demonstration. Have the materials and apparatus arranged neatly and in the sequence of use on the demonstration table. Rehearse the procedure.
2. When the lesson involves materials that may be found in the home, provide an attractive exhibit. For



Do you find Mrs. Goldberg's demonstrations are a little hard to follow?

example, a variety of the acids and bases commonly used in the home include household ammonia, vinegar, boric acid, milk of magnesia, etc. Have students read the labels, particularly those related to safety.

3. Insure good visibility. Use large-scale apparatus, equipment, containers, and exhibits. For example, use large strips of litmus paper which fit into 8-inch test tubes, rather than the small pieces used for individual laboratory work. Provide aids to visibility such as large white cardboards in back of bottles of chlorine.

4. Keep the demonstration as simple and direct as possible. When relatively complex apparatus must be used, take time to have students note and identify important parts. For example, a Hoffman apparatus may look very complicated, particularly to those in the back of the room. Such features as electrodes and their position, the reservoir, the graduations on the sidearm tubes, etc. should be pointed out. The stepwise production of a blackboard diagram as features are located and discussed is well worth the classroom time it requires.

5. Provide for and invite maximum participation by students as demonstrators and observers who report to the class. However, keep in mind that the teacher is the skilled performer. It is not wise to have students perform critical demonstrations for which special training is necessary.

6. When the situation is dramatic, be a good actor. Obvious suspense, enthusiasm, and enjoyment are infectious. Don't be afraid to laugh when the situation calls for laughter.

7. Don't hide behind the demonstration table. Get in front of it and, where possible, move around the room with the material which you wish the class to observe and discuss.

8. Don't "telegraph your punches." Avoid stating the principle or generalization before performing the demonstrations from which it should emerge inductively. One experiment does not support a generalization. For example, if your aim is to show that active metals replace hydrogen in acids, use several different metals and acids before asking your pivotal question. Then try copper and silver with the same acids. The unexpected in a demonstration lends or maintains interest.

9. Emphasize safety techniques. At first, as you demonstrate, call attention to such techniques as proper manipulation of an acid bottle or a test tube. Later ask student demonstrators to indicate the safety procedures before they demonstrate. Occasionally test the class alertness to a violation of safety procedures.

10. Use controls and alert the class to their significance.

11. Sometimes it may be necessary to pass objects around for student inspection. The best procedure is to have a single student take the object around the room.

12. Use occasional failure to good advantage. If a demonstration fails to work in spite of careful preparation, encourage the class to suggest reasons for failure and possible procedures for checking their conclusions. Such exercise in critical thinking provides valuable outgrowths which may outweigh the success or failure of the demonstration.

A series of 64 "Sealed Tube Experiments" is described in detail in the *Journal of Chemical Education.*[9] They have the great value of permanence which permits continuing use and reproducibility.

[9]Campbell, J. A., *Journal of Chemical Education,* 47, 273 (1970).

THE LABORATORY METHOD

There is no doubt that the individual laboratory experiment by the student will continue to be a valuable practice in the teaching of high school chemistry.

Scientific advances have been made by highly skilled experimentors conducting carefully controlled experiments.

Some of the principles of organizing and carrying out a program of individual laboratory work are as follows:

1. In the laboratory students should get first-hand experience with materials to be used in answering a vital question or solving a problem.
2. Sufficient exercises of the acquaintance and drill type should be provided so that the student learns how to handle, assemble, and control materials and apparatus with safety and efficiency before they use them in the laboratory.
3. Proper safety precautions must be observed at all times. One valuable publication which emphasizes safety is *For Greater Safety in Science Teaching*, published by the Board of Education of the City of New York.
4. Laboratory activities should be closely correlated with classroom activity. Laboratory experience should precede classroom application. Ideally, laboratory problems arise from accidental or deliberately engineered situations in the classroom. Students should be encouraged to state a problem and the class should help in making a plan for attacking or solving the problem. The subsequent discussion of the laboratory findings and the resolution of the problem should follow as soon as possible. In this manner the laboratory plays a significant role in training students for scientific problem solving.
5. Powers of observations should be developed and employed. The situation should stress the "how" and the "why" rather than the "what" approach. The observations made should be an honest record which does not reflect the textbook content or what is reputedly "right."
6. Each experiment should be set up as a specific problem to be solved. It should be presented so that students of any ability levels are challenged to do their best. Experiments should be organized so that the simplest ones are performed first and lead into those of greater difficulty.
7. The curriculum should include a number of experiments of the "open-end" type. They encourage the individual project approach.
8. Experiments should be provided which call for the use of related texts, laboratory manuals, references, and handbooks. Adequate library facilities should be provided in the classroom or laboratory. Time should be available for this type of investigation, discussion, and group or individual planning.
9. The teacher's role is that of a guide. He exercises leadership by directing activity. He acts as consultant, critic, and safety engineer.
10. An experiment worth doing is worth recording. The record need not be too formal. The data should be recorded as collected. A minimal report should include the problem, what was done (possibly in outline and diagram), record of observations, conclusions, and reasons for conclusions drawn. Ideally, the report should be written immediately after the experiment has been performed. The lengthened or double laboratory period encourages this desirable practice. The discussion following consideration of reports is the most fruitful part of the experiment.
11. Planning for efficient laboratory work implies some degree of routinization. Students must be assigned regular stations, materials must be ready, and if possible, distributed. The class must have been prepared for the activity. Student aids or laboratory assistants facilitate the efficient use of laboratory time. A regularly scheduled laboratory "day" is sometimes necessary, but ideally careful planning should permit proper timing of laboratory work.

Safety in the School Laboratory[10]

1. Responsibilities of the Instructor: the supervisor of laboratory activities must at all times prevent accidents in such activities.
   a. Teach safety at all times, not just during the first laboratory session.
   b. Circulate among the students during the laboratory period.
   c. Stay in the laboratory.
   d. Emphasize precautions which are specific to specific laboratory experiments.
   e. Plan and/or check the physical facilities to insure maximum safety in their use.
2. Physical Facilities for the Safe Laboratory.
   a. Adequate first aid equipment backed up by a knowledge of what to do in the event of injury.
   b. Appropriate fire extinguishers.
   c. Fire protection for flammable chemicals.
   d. Two exist from the laboratory, as far apart as possible.
   e. Properly designed ventilation systems.
   f. Adequate space for both storage and laboratory work.
   g. Safe methods for the disposal of dangerous waste chemicals and materials.
   h. Deluge showers, if deemed necessary.
   i. All electrical equipment checked out against shock and arcing hazards.
3. Obligations of the Student.
   a. Follow instructions, printed and verbal.
   b. Plan the laboratory experiments to minimize the possibility of accidents.
   c. Attempt no experiments beyond the directed ones.
   d. Become familiar with the health and safety hazards of all equipment and chemicals used.
   e. Record and investigate with the teacher any accidents or unplanned events which occur.
   f. Wear goggles at all times to insure maximum eye protection.
4. Types of Avoidable Accidents.
   a. The most common injury is cut hands, caused by broken glass tubing, other broken glassware, and sharp hand tools.
   b. Chemical spills and uncontrolled reactions.
   c. Splashing hot liquids.
   d. Careless pipetting.
   e. Exposure to vapors or gases.
   f. Foreign bodies in the eye.
   g. Electrical shocks.

   h. Careless handling of mechanical equipment.
   i. Falls.
5. Explosion hazards: avoid the following combinations among chemicals commonly found in the high school laboratory.
   a. Sodium or potassium with water.
   b. Ammonium nitrate, zinc powder, and a small amount of water.
   c. Potassium nitrate with sodium acetate.
   d. Nitrate and an ester.
   e. Peroxides with magnesium, zinc, or aluminum.
   f. Chlorate and sulfuric acid.
   g. Nitric acid with zinc, magnesium, or other metals.
   h. Halogen and ammonia.
   i. Phosphorus with nitric acid, a nitrate, or a chlorate.
   j. Mercury(II) oxide with sulfur.
6. Fire Hazards.

[10] Most of this material is reprinted with permission from *Current Science/Science and Math Weekly Bulletin 15*, a publication of American Education Publications, 1250 Fairwood Ave., Columbus, Ohio.

Some Incompatible Chemicals Found in High School Laboratories	
Chemical	Should not come in contact with
alkali metals; e.g., sodium or potassium	carbon tetrachloride, carbon dioxide, water
alkaline metals; e.g., powdered aluminum or magnesium	carbon tetrachloride or other chlorinated hydrocarbons, halogens, carbon dioxide
acetic acid	nitric acid, peroxides, permanganates, ethylene glycol, hydroxyl compounds
ammonium nitrate	acids, flammable liquids, metal powders, sulfur, chlorates, any finely divided organic or combustible substance
bromine, chlorine	ammonia, petroleum gases, hydrogen, sodium, benzene, finely divided metals
chlorates	ammonium salts, acids, metal powders, sulfur, any finely divided organic or combustible substance
hydrogen peroxide	most metals and their salts, alcohols, organic substances, any flammable substance
hydrogen sulfide	oxidizing gases, fuming nitric acid
hydrocarbons; e.g., propane, benzene, gasoline, etc.	fluorine, chlorine, bromine, sodium peroxide
iodine	acetylene, ammonia, hydrogen
mercury	acetylene, ammonia
nitric acid (conc.)	acetic acid, hydrogen sulfide, flammable liquids and gases
potassium chlorate	sulfuric and other acids, any organic substance
potassium permanganate	sulfuric acid, glycerol, ethylene glycol
sodium peroxide	ethanol or methanol, glacial acetic acid, carbon disulfide, glycerol, ethylene glycol, ethyl acetate
sulfuric acid	potassium chlorate, potassium perchlorate, potassium permanganate, similar compounds of other light metals

7. Safe Use of Radioisotopes[11]

7.1 Possession limit of exempt quantities: See Table.

a. A laboratory may possess up to ten such exempt quantities at one time.

b. Up to ten separate bottles of the same radioisotope may be in the laboratory at one time. These may not be mixed.

c. No restriction exists which limits the number of such scheduled quantities *used* in the course of a year.

7.2 TABLES: Popular Radioisotopes

Radioisotope and Types of Radiation	License-exempt quantity (microcuries) Liquid	Sealed	Half-Life
carbon-14 (β)	50	50	5700 yrs
cesium-137 (β, γ)	1	10	30 yrs
chromium-51 (γ)	50	50	27 days
cobalt-60 (β, γ)	1	10	5.2 yrs
iodine-131 (β, γ)	10	10	8.1 days
*lead-210, RaDEF (α, β, γ)	....	....	22 yrs
phosphorus-32 (β)	10	10	14.3 days
polonium-210 (α)	0.1	1	138 days
sodium-22 (β^+, γ)	10	10	2.6 yrs
thallium-204 (β)	50	50	4 yrs
†uranium nitrate U-238 (α, β, γ)	....	....	4.5×10^9 yrs

*Also natural product, but state laws vary. Available from 0.1 μc to 5 μc depending upon supplier.
†Natural product, no limitation of consequence unless state law to contrary.

7.3 Specific Activity

a. Radiochemicals are usually supplied in aqueous solution in high or low specific activity, i.e. 10 microcuries of I^{131} in 0.5 ml of solution-high specific activity, while the same amount in 5 or 10 ml of solution would be of low specific activity.

b. Dispensing solutions of high specific activity: Micropipets of Lang-Levy type used in conjunction with the Clay-Adams micropipet control. Volume of liquid admitted or expelled easily controlled by twisting the screw at the top with thumb and index finger.

c. Dispensing low specific activity solutions, in the 5-ml range: Use an ordinary laboratory pipet and a remote control device, the Propipetter (hollow rubber sphere with ball valves). For approximate volumes the ordinary medicine-dropper pipet is acceptable.

d. *Under no circumstances should any radioactive liquid be pipetted by mouth nor placed in the mouth.*

7.4 Handling Radioisotopes

a. All open radioactive liquids should be handled over a tray lined with absorbent paper.

b. Rubber gloves should be worn for preparative work but are rarely used when counting samples in order to prevent contamination of counting equipment. They should be monitored by another person for contamination *while they are still on the hands.*

c. Spills should be wiped immediately with absorbent paper (wear rubber gloves) and wet paper placed in *special container for radioactive waste disposal.*

d. Samples are prepared for counting by adding the reaction system to a small shallow cup called a planchet. After the planchet has been dried, it is a good idea to spray it with a transparent plastic lacquer like Krylon. *Do not use spray on an alpha emitter.*

e. Keep all acids away from I^{131} and C^{14} solutions or else use under hood in good working order. This latter precaution should be taken whenever a radioactive gaseous product is expected.

f. Before leaving the laboratory after radioactive materials have been used, the work area, the hands, and the clothing should be monitored for contamination with a Geiger counter with the instrument set on its lowest scale. *Any count above normal background should be interpreted as evidence of contamination.*

g. The common CD V-700 Geiger counter will not detect alpha particles or soft (weak) beta radiations because of the tube thickness (30 mg/cm^2 density thickness). The weak beta radiations of carbon-14 (0.154 Mev) and sulfur-35 (0.167 Mev) may be easily detected with a thin end-window Geiger tube of 1.4-3.0 mg/cm^2 thickness.

[11] Used by permission of John W. Sulcoski (author) and *Science and Math Weekly,* a publication of American Education Publications, Inc., 1250 Fairwood Ave., Columbus, Ohio.

h. Segregate glassware used with radioactive materials and identify with a small dab of paint or nail polish. *AEC approved warning labels should be affixed to glassware containing radioactive material.* The label should indicate the amount of radioisotope inside, the amount of radioactivity in microcuries or counts per minute, and the date of preparation. *Used glassware must be decontaminated before being used again.* Scrub with a powerful laboratory detergent such as Sparkleen, Alconox, Radiowash, or even Tide. Rubber gloves should be used along with plenty of running water as a rinse. For glassware difficult to decontaminate, soak for a few days in chromic acid bath (100 g $K_2Cr_2O_7$ and 400 ml of conc. H_2SO_4 in 600 ml of water). If still contaminated, throw it away. Dried glassware should be monitored with Geiger counter for residual radioactivity which should be no higher than background count.

i. Waste Disposal. All material which cannot be decontaminated, such as paper used in trays, disposable syringes, solid waste, plastic spatulas, etc. should be placed in a storage can lined with paper and specifically marked for disposal of radioactive waste. As often as possible, the waste should be disposed of in small quantities along with ordinary laboratory refuse. Repeated scrubbings with a strong cleaner will usually remove contamination from workbenches, floors, and other surfaces. One laboratory sink, preferably one seldom used, should be designated as a "hot sink" for the disposal of radioactive wastes and rinsings.

j. Storage: Radioactive sources and liquids should be kept segregated in a small box or cabinet *that can be locked securely.* Thick shielding such as lead is unnecessary with license-exempt quantities. A small plastic box or a small tin can will provide adequate shielding. Commercial type lead shields can be purchased or secured gratis from the radioisotope lab of a local hospital. The amount of radiation emitted by license-exempt quantities of radioisotypes does not necessitate the wearing of a film badge or dosimeter.

THE PROJECT METHOD

"A project is an extended problem or series of related problems that are the outgrowth of the student's own interests and endeavors at problem solving."[12]

The project method is particularly suitable for special groups. Honor chemistry classes may use the method almost exclusively. Special project classes may be formed under other circumstances. A chemistry club furnishes another opportunity for individual or group projects.

The following list of suggestions will guide the interested teacher.

1. Seek to locate, suggest, and encourage projects of varying complexity so as to offer problems that will challenge and stimulate a wide range of student abilities and interests.
2. Organize and/or use a science club as a forum for extending and exchanging student experiences with projects.
3. Get started early. Even the first or second week of school is not too early to introduce the idea and practice of projects.
4. Call upon the many resource persons to be found in almost any community—people employed in science-related professions and occupations. Encourage students to participate in Science Summer Institutes.
5. Get ideas from periodicals, reference works, and publications available to chemistry teachers. Among these would be *Chemistry*, *The Science Teacher*, *Scientific American*, *Journal of Chemical Education*, and other periodicals, and the booklets and pamphlets put out by the National Science Teachers Association, the American Chemical Society, and Science Service.
6. Participate in various meetings of science teachers where exchanges of experiences and ideas are an important part of the program.
7. Be patient in establishing the project method. Don't try to move from zero all the way to the boiling point in one grand swoop.
8. Let the students help you. Allow time for students who have done projects to tell others about their work. Tell your students about the project successes and honors gained by students of former years.
9. Continually keep an eye on the types of activities in which students are engaged, lest some drift into unguided dabbling of little educational worth.
10. Students should be encouraged to report their projects outside as well as in the classroom. Reports can be submitted in such competitions as Tomorrow's Scientists and Engineers or the NASA Science Congress both of which are sponsored by the National Science Teachers Association (NSTA). Student project reports are frequently published in *Chemistry* and in the *Science World.*

The project method affords the highest type of training in scientific method. Each student takes responsibility for completing a project individually or as a member of a group. Most project activities do involve some degree of library research, review of current information from a variety of sources, and probably some repetition of previously known experimental laboratory procedures and findings. Projects generally

[12] Miller, David F. and Blaydes, Glenn W., *Methods and Materials for Teaching Biological Sciences,* page 45, McGraw-Hill Book Company, 1938.

require experimentation and other research activities. The classroom or laboratory becomes a workroom with storage facilities, tools, laboratory equipment, and reference books and materials available as needed. Phases of it, or even entire projects, may be carried out in the home workroom or laboratory and in library facilities outside the immediate school environment.

THE DEVELOPMENTAL METHOD (Class Discussion of a Problem)

Fortunately there are few units in chemistry which do not readily lend themselves to the three methods previously discussed. In other cases the main reliance will be on the textbook or books and on suitable reference materials. In a unit of this kind the teacher should use other types of teaching aids to the fullest extent.

The developmental method is indicated for units or subunits dealing with theories such as the atomic theory, the electron theory, and the ionic theory, or for those dealing with the concepts and fundamental skills of formula and equation writing and mathematical problems.

The atomic theory may be developed from the problem, How were the fundamental laws of chemistry explained? In approaching the problem, the class would consider the birth of chemistry as a modern science and the developments from 1775-1805 which challenged scientists to explain the nature of matter. Discussion of the roles of Democritus, Lucretius, etc., in leading to Dalton's organization of ideas into a conceptual scheme would follow. The prediction of the law of multiple proportions and its verification, the concepts of atom and atomic weight, molecule and molecular weight, and the language of chemistry (symbols and formulas) would then be considered as outgrowths of Dalton's conceptual scheme.

The emphasis in such lessons is not on presentation but rather on inductive development. The student does not recite but rather recognizes the relationships between facts and states the concepts as a result of classroom experiences. The teacher notes for each chapter give suggestions that can be used for this method.

The principles of the method include the following:

1. The interest of the class should be aroused by a method of introducing the problem which makes students recognize a need for the knowledge.
2. Motivation should lead directly to the aim of the lesson, which should be stated by the students, written on the board, and kept in view constantly throughout the lesson.
3. Problem situations should be presented for consideration. These should require ability and give practice in logical thinking.
4. The teacher should use analytic questioning to guide the class to a solution of the problems.
5. A minimum of information should be given by the teacher and a maximum drawn from the students.
6. As much concrete and illustrative material as possible should be used.
7. Facts should be clearly differentiated from inferences and explanations.
8. A chalkboard outline should be developed as the lesson unfolds.
9. Medial summaries should be used at intervals and a terminal summary should provide the final generalizations.
10. Application should test understanding of generalizations.
11. Whenever possible, the assignment should emerge during the lesson as a part of the current problem.
12. Teacher domination of the lesson is to be avoided.

The principles given above apply not only to the developmental lesson, but to other lesson types as well.

There are many stimulating variants of the developmental method. Socialization techniques may be used, in which students carry on the discussion under the unobtrusive direction of the teacher. The presentation of selected student reports, followed by a question-answer period, is one such device. The deadly practice of having students read lengthy reports should be avoided. Reports should be presented orally, with the aid of prepared outlines. Such reports should be supplemented by whatever illustrative materials the student and teacher think appropriate. Debates, dramatizations, or other programs using the format and technique of current popular TV and radio presentations are other devices which stimulate and maintain student interest.

OTHER METHODS

The *review* or *drill lesson*, particularly before a unit examination or in preparation for midterm or final examinations, is a necessary and valuable instrument of instruction. Standard practice employs the recitation, question-answer type of lesson. This may take the form of review of past examination questions. The procedure permits maximum use of time, complete and organized coverage of subject matter, board work, and the use of teaching aids.

The review lesson need not be confined to examination questions. Some effective methods that may be employed include:

1. Student-suggested questions as the basis of the review.
2. A spelling-bee type of lesson, with the class organized into teams. This may be effective with definitions, equations, principles, formulas, etc.
3. A board of student experts, each of whom in succession is asked questions by individual students. Each student has the chance to ask one or more questions previously prepared.

4. Procedures based on quiz programs in vogue on TV. They provide for individual differences but require that the entire class play a part by providing questions, alternate contestants, etc.
5. A film which covers the unit. The value of this depends on the excellence of the film and the activities which precede or follow its showing.

The *lesson which is organized around the use* of a film or some other type of audiovisual aid merits special consideration. This lesson provides the student with a background of experience when direct contact with actual materials through demonstration or laboratory experiment is not practical. Such aids to instruction should be carefully appraised before use. Where the direct approach does the job just as effectively, a film or filmstrip has little or no value. But there are many unit areas in which such materials may be indispensable. Carefully selected films will be suggested for use with specific lesson units in the teacher notes. Some of these will be invaluable as previews of units or as summaries.

To insure best results when using audio-visual aids, the following procedures should be employed:
1. Preview the material and prepare an introductory discussion to show its relation to the unit being taught. If a teacher's guide for a film is available, use it to plan this discussion. Prepare a lesson plan and make the material an integral part of this plan.
2. Prepare the class before presenting the film or other aid. This may involve not only the preliminary discussion, introduction, or problem, but also mimeographed aids in directing student attention to special parts of the film. These may be in the form of questions to be answered orally, or in writing.
3. Use films and other such aids to supplement, not supplant, the teacher and other teaching procedures.
4. Carefully schedule such materials that must be borrowed so that they fit into the unit sequence.
5. Vary the use of the audio-visual aids which are available. Always select the best method and the best materials. For example, the formation of crystals from melt or solution may require the use of the microprojector.
6. Provide for definite follow-up activities. These may be written questions to be answered, a written report, a discussion period, subsequent related library investigations, etc.

The use of some of the audiovisual aids requires special training. The teacher operating film devices must be thoroughly trained in their effective use, and in many cases, their proper maintenance. A number of technical difficulties may interfere with the best use of any film. At the start of class the visual material should be ready including any accompanying audio material. The timing of the presentation should have been checked so that it will not run over class time.

The *field trip* to an industrial plant or laboratory has special appeal and teaching value. It enriches the course by stimulating student interest and uses the community resources to correlate classroom work with real experience. Field trips may include visits to sewage disposal plants, water purification plants, food plants, mines and quarries, engineering projects, and oil refineries. Trips may also be taken to local science fairs, industrial fairs, the science "Open House" at local universities, and science "Career Conferences." The latter activities may be carried out on an individual as well as a group basis.

The teacher must explore the place to be visited and arrange the details. These include the proper clearance and scheduling with parents, school authorities, and transportation companies, as well as with the public relations representatives of the place to be visited. Preliminary briefing sessions and supplementary lectures may be required. A mimeographed guide and direction sheet is frequently helpful. It may include an outline map of the plant with notes on objects especially worthy of observation. Questions may be listed as a basis for field work and for later class discussion.

The following procedures are helpful in carrying through a field trip:
1. Inform pupils of the amount of money they will need for carfare and lunch.
2. Require each student to submit a parent's consent slip in advance.
3. List the address and telephone number of each student for use in case of emergency.
4. If the students will miss any classes during the trip, advise the teachers concerned, several days ahead of time.
5. Do not take too large a group. Occasionally a small representative committee from a number of classes may be selected to go on a trip.
6. Provide for adequate supervision. Students, particularly those who are members of clubs or squads, may act as leaders and field assistants. They should be briefed in advance. Student teachers and parents may be enlisted.
7. See that each member of the group is an active participant.
8. Follow up the trip with discussions, reports, exhibits, and laboratory study.

Field trips help develop understanding of the role of the producer and consumer. A student sees the careful inspection procedures, the rejection of products not up to standards, and the testing procedures employed by industry.

Field trips encourage understanding of the roles of various levels—professional, technical, and untrained labor—in productivity, and appreciation of the complexity and correlation of effort involved.

Another method or plan of instruction that should be considered is the *committee method.* The *committee method* is an effective device for cooperative planning and learning. It may be valuable in units dealing with

theories, nuclear energy, and metallurgy. A rotating committee may be used in dealing with current events. The various problems to be studied are decided upon and assigned or selected. Group leaders coordinate, plan, and are responsible for group accomplishment. The groups then solve their assigned problems separately. When necessary, progress reports are submitted to the group and further directions of inquiry are selected. When the work has been completed, the committee reports are given in one or more class meetings. In the reports, materials are exhibited and any experiments developed in the course of the group work are demonstrated or described. Questions from the group are encouraged. During the group activity, minimum requirements (as in basic readings) are met by all students. The class as a whole should make some evaluation of the unit of work presented by the committee. A test should then conclude the unit.

NEWER METHODS

Modular scheduling has caused a shift toward *individualized instruction.* One experiment describes the division of the science day into 15-minute modules with two or more of these devoted to (1) teacher-directed classes (up to five times a week); (2) problem day; (3) question day; (4) open laboratory sessions—student selection from among open-ended laboratory investigations.

The *individual method* in its various forms attempts to provide for individual differences by permitting each student to progress at his own rate. This scheme meets one particular problem of instruction but does not offer the socialization values of the more conventional methods of instruction. The method has been modified in its most successful use. All members of a class group study a particular unit, but those who progress most rapidly do more supplementary work. Upon completion of the unit within a set time limit, the individual experiences and learnings are collected, correlated, and summarized by the group as a whole.

Newer techniques also include team teaching, programmed instruction, and teaching by film. The team teaching approach generally involves a teacher specialist in specific areas presenting the material to large student groups. Suitable demonstrations are utilized along with other aids such as films or transparencies. The lecture or lecture-demonstration method is frequently employed. The group is then split into smaller sections for recitation, explanation, further exposition, and quiz. Representative programs available are listed in the Bibliography.

Ms. Buzini is the only teacher in our school who has time enough to do everything.

Part II

Teaching Suggestions for a Modified Chemistry Course

Since courses in *Modified* or *Applied Chemistry* vary considerably in content and sequence, the suggestions in this part represent only one of many possible plans. A fine outline of a modified chemistry course may be obtained from the Bureau of Science Education, The University of the State of New York, Albany, N.Y. 12224. The course of study for General Chemistry 1970 was developed during several years of experimentation. It is mainly for the average student whose interests and goals may be different from those who take the standard high school chemistry program. The committee who produced the course consisted of experienced high school teachers, representatives of two-year Community College, and representatives of schools of Nursing. *Modern Chemistry 1974* may be used as the basic text for such a modified course.

The Teacher's Notes in Part VII may be used continuously as a basic reference for activities, demonstrations, teaching aids, experiments, exercises, and tests. The standard program is changed for the modified courses. The experiments and exercises mentioned are from the accompanying laboratory manual to *Modern Chemistry 1974, Exercises and Experiments in Chemistry.* Items from *Tests in Chemistry* and/or *Alternate Tests in Chemistry* may be selected to evaluate the topics chosen. For example, to promote reading comprehension, portions of the text may be read aloud. Then the students may be assigned to answer the accompanying exercise. The exercises are arranged in a sequential manner to accompany the text. The answers to the exercises may be reviewed orally in class. This method can pinpoint student difficulties before they become problems.

Desirable laboratory skills and techniques are developed and strengthened by the use of selected experiments from among the experiments in *Laboratory Experiments in Chemistry.* Many of the qualitative and quantitative experiments may be used to provide basic training for the students who are interested in chemistry-related technological careers.

UNIT 1
CHEMISTRY IN A MODERN WORLD

This unit, which serves as an introduction to the study of chemistry and applied chemistry, includes the conventional concepts usually developed. In Chapter I, The Science of Chemistry, Sec. 2.15 may be omitted. In Chapter 2, Matter and Its Changes, you may wish to omit Sec. 1.17-1.21. You may wish to change the sequence of sections. One sequence may be:

1. What is Chemistry?
2. Why do we study chemistry (applied chemistry)?
3. What is matter?
   a. What are the properties of matter and how are they recognized or ascertained?
   b. What are the phases of matter?
   c. What are the classes of matter?
4. What changes may matter undergo?
5. What are the methods which scientists employ in solving problems?
6. What are some of the fundamental laws of chemistry?

It is suggested that early excursions into demonstrated physical and chemical changes be made in order to stimulate and maintain student interest. Considerable time should be devoted to the role of chemistry in students' everyday living, in the economic

and technological aspects of industry, in the development and conservation of national resources, in leisure time activities and hobbies, and of course, in the career possibilities related to the teacher level of ability and interest.

Before proceeding to the next unit, it is frequently worthwhile to devote some time to a review of the basics of the chemistry of oxygen, hydrogen, and water. It will provide a feeling of reality and relationship to more-or-less familiar aspects of chemistry. The review should be laboratory-centered by having students perform experiments dealing with oxygen, hydrogen, and water. These include Exps. 5, 14, 15, 18, and 19 in *Laboratory Experiments in Chemistry.* The laboratory work will provide the students with an understanding of chemical changes. The atmosphere of the laboratory should support the idea that chemistry is an experimental science.

UNIT 2
THE ORGANIZATION OF CHEMISTRY

Chapter 3
ATOMIC STRUCTURE

Diagrams of the nuclei of the first 20 elements and their isotopes may be made by the students. The material in Sec. 3.12 and Sec. 3.13 may be too difficult for certain classes, but it is worthwhile to make the attempt to expose them to the related ideas. The use of a carton of a dozen eggs, or a packaged ream of paper relates the idea of functional units in everyday life to the idea of the unit in chemistry, the mole.

Sometime during this or the next chapter, good use may be made of the following "black-box" demonstration-experiment. The teacher exhibits a closed box (a shoe box will do) and asks students to find out what is inside, without opening the box. After a discussion of how to find out, students manipulate the box and generally come up with the idea that a small solid object is inside. The relationship of the demonstration to the problem of atomic structure is established by class discussion. Such points as investigation by observation of macroscopic properties and manipulation, the use of apparatus or instruments, formulation of hypothesis-theory or model, etc. are made. The entire procedure helps students appreciate how our knowledge and model(s) of atomic structure evolved even though an atom is too small to be seen.

Chapter 4
ELECTRON CONFIGURATION OF ATOMS

Much of this chapter is omitted: Sec. 4.1-4.5 and Sec. 4.7-4.9. However the teacher will find it useful to develop electron-dot symbols in Sec. 4.6 because of their utility in bonding in Chapter 6. Some classes can be taught the idea of sublevels, and both the electron configuration and orbital-notation schemes are extremely helpful as they relate such electron-dot symbols as that of carbon. The *s* and *p* sublevels and the corresponding schematics may be readily mastered by the average class.

Although the formal treatment of spectra and quantum numbers may be omitted, students can be taught the idea of excited electrons (atoms) and the emission of characteristic radiation by return of the electrons to lower energy levels. Students may perform Exp. 45, "Flame Tests," in conjunction with the demonstration of spectral lines and the spectroscope.

Chapter 5
THE PERIODIC LAW

Probably the best and simplest approach is to start with student work sheets which may have been made in connection with the previous work on Chapters 3 and 4. Such work sheets have the first 20 elements arranged in periodic fashion according to atomic number (and mass numbers, if desired). Each box has an atom model and the lower right may contain the representative electron-dot symbol. Students may fill in other data in the lower lefthand corner as it is developed in the text. Students are asked to indicate how elements of the last vertical row (Group VIII) are alike in structure. This emphasizes the octet and electron-pair (He). Brief discussion of the uses of helium and possibly neon and argon leads to the idea of lack of chemical reactivity as related to the existence of the octet. You may tell the story of the discovery of the compounds of xenon and krypton. Then ask students to predict chemical reactivity of the elements in the other Groups. Have them find out how the elements of Groups I, II, and III may attain an octet. Have the class determine what kind of elements aluminum and magnesium are. This leads to a definition of a metal on the basis of structure. Treat groups VII and VI (V and IV also, if desired) in the same way, and end with a definition of a nonmetal. Have students explain the meanings of group, family, and period. Have them discuss variation in properties across a period and down various groups. Variations of atom size are readily related to atom models and the number of main energy levels. Most students easily grasp the idea of ionization energy and the ways in which it varies in a group and across a period. Details associated with Sec. 5.7 may be omitted entirely or merely mentioned. Sec. 5.8 may be treated in the same manner.

UNIT 3
CHEMICAL FORMULAS AND EQUATIONS

Chapter 6
CHEMICAL BONDS

The use of atom-shell and electron-dot-diagrams helps students master the details of bonding. The orbital notation model should be used sparingly since it may pose difficulties for many students. In the first lesson use examples, primarily of the AX type, which students work out. Focus attention on the principle of charge equality.

In Lesson 2, expand the treatment to the AX_2 and A_2X compounds. Students can master the first three oxidation rules by slowly proceeding from their experimental experiences with oxygen and oxidation involving oxygen. Burn sodium in chlorine and in air. Burn magnesium in air. It is advisable to underemphasize oxidizing agent and reducing agent concepts at this particular time. It may be necessary to spend an extra lesson on the oxidation rules. Comparative sizes of atoms and ions are developed from the atom model diagrams and may be related to Table 6-3.

A final lesson on ionic bonding should be used to review previous examples and concepts. Students will feel secure if a simple rule dealing with combinations of elements of Group I and II with those of Group VII or VI is discovered. The lesson may also deal with significant properties of ionic compounds, i.e. crystalline solids of high melting point are generally water soluble. These properties should be demonstrated, i.e. melt ice and attempt to melt NaCl. The first lesson on covalent bonding should deal with *pure* covalent (nonpolar) bonds. Use H_2 and F_2 as an introduction and illustrate with atom-model and electron-dot structures. Develop the idea of a molecule and molecular formula. Then use Cl_2, O_2, and even N_2. Oxygen and N_2 introduce the concept of multiple bonds. Treat the energy and mole concepts lightly, if at all.

Devote the second lesson to the polar covalent bond concentrating on HF, HCL, H_2O, and possibly NH_3. Develop oxidation rules for hydrogen and oxygen. Use Exp. 8 "Covalent Molecules." Emphasize the concept of the molecule and its relation to structure. You may wish to develop the idea of a polar versus a nonpolar molecule on the basis of examples from these first lessons. A third lesson deals with electronegativity as an operational basis for developing oxidation Rule 6, dealing with bond polarity. This lesson concentrates on Sec. 6.15-6.16. You may wish to deal with CH_4 and CCl_4 molecules as examples of structures which have polar bonds but nonpolar molecules. Students may discover the variations in electronegativities by detailed study of the table.

A fourth lesson concentrates on the content of Sec. 6.13. This deals with simple formulas of molecules. Monatomic and diatomic molecules are emphasized. The more complicated molecules such as O_3, P_4, and S_8 may be mentioned but you will probably find it advisable to treat the symbols only for phosphorus and sulfur in future equation work. Students need drill on the meaning and use of coefficients.

You may wish to use portions of a last lesson to have students understand how a polyatomic ion is structured. Use simple examples such as ammonium, hydronium, and hydroxide. This treatment gives the students some understanding of what a radical is, how it is represented, and possibly oxidation Rule 7.

Chapter 7
CHEMICAL COMPOSITION

The primary aim of this chapter is the development of the fundamental skills in formula writing. The suggested procedural sequence is development of understanding, learning the rules, application and drill, and daily quiz. See quizzes in Answers to Questions and Problems Section for Chapter 7. Quizzes may be marked by student interchange of papers immediately following the quiz. In this way sources of error are apparent and are immediately followed by relearning of correct procedures and application. As a start, use Exercise Sheet #1 in Chapter 7. Have the students develop the charges on these ions from the basic formulas given, i.e. compounds of sodium and chlorine. Simultaneously have them learn the names by transference of determined charge from the "given" formulas to the place provided. Then have them apply the stated rules to a few simple examples, i.e. CuO, Na_2S, $Ca(OH)_2$, $Al_2(SO_4)_3$, etc. in class. Schedule a quiz for the next day on the ions and charges appearing in the exercise and assign the task of completing the formulas in the exercise. On the day following the quiz send students to the board with blank sheets to complete the vertical columns. Have the class correct boardwork and homework immediately thereafter. During the boardwork you may extend the tabulated ions to include others selected by giving additional formulas, i.e. sodium sulfite, sodium hydrogen carbonate, sodium bromide, etc. Fold over the right side of the exercise sheet as far as the first column. This

procedure permits the class to complete formulas for the "new" anions. You may wish to extend the procedure to other cations, i.e. Cr^{+3} or Ni^{++}. Schedule a quiz for the next day on writing formulas. It will probably be necessary to schedule a second quiz on formulas for a subsequent day. A scheme which works for students who are having trouble is to permit them to take a test (quiz) within several days, at their request. This procedures saves many students who are ready to give up when they do not score well on the regular quizzes. See quizzes, pages 38 and 39.

Chapter 8
CHEMICAL EQUATIONS

Introduce this chapter by using Exp. 10, "Balancing Chemical Equations." This experiment emphasizes the law of conservation of atoms. It permits students to manipulate models in balancing equations involving molecular substances. It also permits review of the electrolysis of water.

Devote the next lesson to Sec. 8.1-8.3 with appropriate drill.

A third lesson may use Exp. 12, "Types of Chemical Reactions." This lesson gives students experiences which enable them to appreciate that many chemical reactions fall into one of these categories. Recognition of the kind(s) of reactant substances permits an "educated guess" as to the resultant products.

You may wish to introduce students to the *mole concept* as treated in Sec. 8.10 and the mole relations involved in a balanced equation. Some students solve mass-mass problems more easily by using the proportion method instead of the mole method.

UNIT 4
THE PHYSICAL STATES OF MATTER

The unifying theme of this unit is the essentials of the kinetic theory as they relate to the three states of matter. Most courses will emphasize the subject of water as an important compound. It is suggested that this phase of the unit treatment be extended to include selected sections of Chapter 12, and the role of water as a solvent.

Chapter 9
THE GAS LAWS

Two lessons are suggested. The first lesson includes Sec. 9.1-9.4. The second lesson deals with Boyle's and Charles' laws qualitatively and includes Secs. 9.5-9.12. Problem work should be minimal.

Chapter 10
MOLECULAR COMPOSITION OF GASES

Use Sec. 10.1-10.4 to help students appreciate the significance of coefficients in equations involving gases. Develop understanding of the Avogadro principle and the diatomic nature of elementary gas molecules. A second lesson is devoted to review and application to gas volume-gas volume problems, Sec. 10.9. This may be extended to involve volumes of air required by treatment of air as approximately one-fifth oxygen by volume.

Chapter 11
LIQUIDS-SOLIDS-WATER

The first lesson concentrates on the physical properties of liquids and gas-liquid phase changes. Related demonstrations include vapor pressure demonstrations, effect of evaporation of various liquids on temperature, saturated cotton rolled about thermometer bulb, boiling under reduced pressure (boiling by cooling), etc.

A second lesson includes Sec. 11.9-11.11 treating solids in the same way as liquids were treated in the first lesson.

The third lesson on crystals, Sec. 11.13-11.14 should include demonstrations on the various methods of crystal formation: evaporation from solution, cooling a molten mass (monoclinic sulfur), sublimation (iodine), and cooling a hot concentrated solution of a very soluble solid such as ammonium chloride or potassium nitrate. Omit crystal systems as required but have students observe cubic salt crystals, two types of sulfur crystals, etc. under hand lens. Exp. 22, "The Solid State, Crystals and Crystallization," may be performed.

The first lesson on water covers Sec. 11.15-11.16, a review of the structure and properties of water molecules. The concept of the hydrogen bond is one that many students can understand. The second lesson on the chemical properties of water includes Sec. 11.17 and Exp. 17, "Chemical Properties of Water." It serves as an introduction to the subject of acids and bases and provides opportunity for drill in writing simple equations.

UNIT 5
SOLUTIONS AND SUSPENSIONS

Chapter 12
THE SOLUTION PROCESS

The first lesson may be spent in demonstrating and contrasting solutions versus coarse suspensions: clarity, settling, filtration, size of particle, etc. Various types of solutions are also demonstrated as well as miscibility and nonmiscibility. The colloidal state may be treated briefly. You may wish to demonstrate the Tyndall effect, Brownian movement, filtration, settling, particle size, and coagulation. Some useful demonstrations are:
1. Dilute India ink 1:5. Add a few drops of acetic acid to a 5 ml sample and filter.
2. Use the Instructor Experiment in Exp. 30, "Adsorption."

The second lesson concentrates on Sec. 12.5, 12.9, and 12.11. and kinds of solutions.

The third lesson concentrates on the effect of temperature and pressure variations on solubility and interreaction plus some simple problems on solubility curves. Schedule Exp. 22, "Solubility, Rate of Solution, Heat of Solution," and Exp. 23, "A Solubility Curve." The heat of solution, boiling and freezing point effects, and selectivity of solvents should be treated from the practical, functional viewpoint. Students may perform Exp. 20, "Solution and Molecular Polarity," as an aid to understanding why "like dissolves like."
3. *Starch*: To simulate a colloid mill (mechanical dispersion), prepare a suspension of starch. Filter, and test filtrate with iodine solution. Grind another starch sample vigorously in a mortar. Add water, filter, and test the filtrate with iodine solution. Try heating a dilute suspension of starch and note change to opalescent colloid. Filter and repeat iodine test.

From this point, the sequence and selection of topics may vary considerably. It is probably advisable to defer Chapters 13, 14, and 15 to a later date because the related equations for ionization-dissociation, and acid-base reactions involve skills that differ from those used in writing formula equations. Some teachers may choose to make an intensive study of a particular element and possibly the Family of which it is a part. Thus Chapter 16 "Carbon and Its Oxides" may be the choice of one teaching schedule. Other teachers may choose the halogens first because the large number of available demonstrations and experiments serve to maintain student interest. One sequence of topics which might be employed is the following:

Chapter 29 "The halogen Family"
Chapter 23 "The Metals of Group I"
Chapter 24 "The Metals of Group II"
Chapter 28 "Sulfur and Its Compounds"
Chapter 27 "Nitrogen and Its Compounds"
Chapter 16 "Carbon and Its Oxides" (Completing the nonmetals)
Chapter 26 "Aluminum and the Metalloids" (with emphasis on aluminum)
Chapter 22 "Elements of Period Three"
Chapter 13 "Ionization"
Chapter 14 "Acids-Bases-Salts"
Chapter 15 "Acid-Base Titration and pH"
Chapter 26 "The Transition Metals"
Chapter 17 "Hydrocarbons"
Chapter 18 "Hydrocarbon Substitution Products"
Chapter 30 "Radioactivity"

The Chapters omitted in the sequence are:
Chapter 19 "Chemical Kinetics," which might be omitted almost in entirety.
Chapter 20 "Chemical Equilibrium," most parts of this chapter could be omitted with the possible exception of the sections which deal with reactions that run to completion, and hydrolysis of salts.
Chapter 21 "Oxidation-Reduction Reactions," the extent to which this chapter is used will vary considerably. Certainly some use must be made of the relative strengths of oxidizing and reducing agents, particularly as this relates to the displacement reactions of metals and nonmetals. Since electrolytic processes have important industrial applications, the topics of electrolysis and electroplating should be included. Student interest in batteries calls for some brief consideration of electrochemical cells.

Whatever the sequence of chapters selected, the order of teacher's materials is in order chapter-by-chapter.

Chapter 13
IONIZATION

The first lesson is a demonstration distinguishing between electrolytes and nonelectrolytes using a conductivity apparatus. Show that a metal completes the circuit in the conductivity apparatus and that a nonconductor does not. Use the metal blade and the wooden handle of a spatula for this procedure. On the board list the electrolytes and nonelectrolyte and by name and formula and point out that electrolytes are acids, bases, and salts.

Demonstrate the rise in boiling point of a salt solution (NaCl). Indicate that the effect of electrolytes on freezing point depression is larger than that of nonelectrolytes for solutions of equal concentration.

This effect is also obvious on the boiling point elevation for nonvolatile substances. Briefly review the ionic structure of NaCl. The ions become free-moving in water solution hence complete the circuit and conduct the electricity. Write a simple equation:

$$NaCl \rightarrow Na^+ + Cl^-$$
(free-moving)

To avoid having students reach the incorrect conclusion that the number of positive ions equals the number of negative ions, have them attempt writing a similar equation for the dissociation of a compound such as $MgCl_2$. Have them also explain why nonelectrolyte solutions such as alcohol or sugar do *not* conduct. (They do not exist as ions in solution.) You may also wish to extend discussion to their nonpolar nature (Sec. 13.9).

The second lesson should be largely devoted to developing mastery in writing dissociation and ionization equations. Start with salts, then the strong and relatively strong bases (ionic) and develop the idea that the charge on the ion is the same as its oxidation number. It is probably advisable to avoid the use of the postscripts (s) and (aq) and possibly after one or two examples to leave out the initial charges on the ions in the solid phase. Then switch to some simple acids such as HCl and show ionization in simple form. This may also be used to show ionization of NH_3. The general outline suggested is:

Dissociation

Salts

$NaCl \rightarrow Na^+ + Cl^-$

$CaCl_2 \rightarrow Ca^{++} + 2Cl^-$

$Na_2SO_4 \rightarrow 2Na^+ + SO_4^=$

$Al_2(SO_4)_3 \rightarrow 2Al^{+++} + 3SO_4^{--}$

Ionic Hydroxides (Bases)

$NaOH \rightarrow Na^+ + OH^-$

$Ba(OH)_2 \rightarrow Ba^{++} + 2OH^-$

Ionization

(Acids)

$H_2O + HCl \rightarrow H_3O^+ + Cl^-$

$2H_2O + H_2SO_4 \rightarrow 2H_3O^+ + SO_4^{--}$

. . . . . .

(Base)

$NH_3 + H_2O \rightarrow NH_4^+ + OH^-$

For acids, you may wish to then use the simple equivalent:

$HCl \rightarrow H^+ + Cl^-$

$H_2SO_4 \rightarrow 2H^+ + SO_4^{--}$

To test student understanding of significance of equations, ask questions such as: How many ions result from dissociation of one formula unit of $Al_2(SO_4)_3$? Or for classes who have dealt with mole concept: How many moles of ions result from the dissociation of one mole of $Al_2(SO_4)_3$? Develop the idea that polar substances (acids and ammonia) ionize by reacting with the solvent, water (polar). You may wish to demonstrate precipitation of slightly soluble salts from mixed solutions and write the ionic and net equations.

Start a third lesson with a review of equation writing. Devote the major part of the lesson to strength of electrolytes, Sec. 13.7 and 13.11. Use glacial acetic acid with conductivity apparatus and have students predict behavior. Then add water and stir and have them note effect on conductivity bulb. Repeat demonstrations using the hydroxides of Group II and contrast with those of Group I. Also use NH_3 solution. Have students note intensity of light as an indication of strength and degree of ionization. Write equation for ionization of acetic acid. You may wish to use very simple equations for ionization of two such acids as HCl and H_3BO_3. Students may incorrectly predict that the H_3BO_3 will be a very strong electrolyte conductor since a simple equation could be $H_3BO_3 \rightarrow 3H^+ + BO_3^{-3}$. Test with conductivity apparatus and finally substitute an ammeter which will give a positive reading. They generally get the idea that the degree of ionization is responsible for the strength or weakness of various acids, particularly when it is brought to their attention that H_3BO_3 is boric acid, an eyewash.

Devote the last lesson to electrolysis (Chapter 21). Start with electrolysis of a concentrated solution of hydrochloric acid using a U-tube and carbon electrodes. The products, hydrogen and chlorine are easily identified.

Repeat with electrolysis of a solution of $CuCl_2$ with results again easily demonstrable. Finally demonstrate water electrolysis and treat as in the text. This is related to Exp. 9, "Balancing Chemical Equations" and another version is described in Further Experimentation, Exp. 43B, "Electrolytic Cells." This lesson series reveals the behavior of ions during electrolysis. You may also wish to develop the principles of electroplating as an application of the ideas and demonstrate copper plating.

Chapter 14
ACIDS, BASES, SALTS

For each class of compound use the same general plan:

1. Exhibit numerous samples related to student environment and usefulness in course.
2. Write formulas, chemical names, and, where appropriate, common names.

3. Demonstrate properties: use appropriate equations. Safety factors.
4. Nomenclature and classification as in acids: binary, ternary.
5. Methods of preparation: general, selected.

Lesson one on acids follows this pattern: Exhibit fruits, vinegar, etc., and industrial acids. Use large bottles in which the latter come to the laboratory. Test samples with litmus paper. From formulas and indicator behavior have students frame Arrhenius' definition. Demonstrate properties 3, 4, 6 in Sec. 14.5. Initial definition along the lines of 1 in Sec. 14.5. Possibly write simple equations: $HCl \rightarrow H^+ Cl^-$. This lesson uses Sec. 14.1, 14.2, 14.3, and parts of 14.5. You may wish to demonstrate safety in diluting acids (H_2SO_4) and the procedures used in pouring acids from bottles.

Lesson two reviews lesson one and is used to develop the modern definition of acids as the course requires. Use Sec. 14.3 and 14.4. Treat sulfuric acid from viewpoint of complete one-stage ionization. Omit multistage ionization of polyprotic acid, i.e. H_3PO_4. Review concept of strength of acids. Schedule Exp. 24A, "Preparation and Properties of Acids." Lesson three focuses on neutralization. Demonstrate Exp. 28, "Titration of an Acid and a Hydroxide, Heat of Neutralization." Test hydroxides with litmus and phenolphthalein. When the demonstration of titration is performed, recover the NaCl by evaporation. Develop the equation in formula and ionic variety. Have students write formula equations for other examples, i.e. $Ca(OH)_2 + HNO_3$, etc. Use Sec. 14.5, part 5. You may wish to show other acid reactions, i.e. with carbonates and oxides.

Lesson four deals with preparation of acids: Exp. 24A, "Preparation and Properties of Acids, and Sec. 14.7 (text), plus a review of preparation of halogen acid, Ch. 29, which would have been previously studied in the suggested sequence. Develop concept of anhydrides. Also use processes of direct union and salt desired acid + sulfuric acid. (Also refer to preparation of HNO_3). Use appropriate equations for various methods.

Lesson five deals with classification and naming of acid, Sec. 14.6. Review methods of preparing acids.

The first lesson on bases parallels the first two on acids. With the preliminary work on acids, this lesson may cover content of Sec. 14.8 and 14.9. Some teachers defer discussion of amphiprotic hydroxides until they reach the topic of aluminum or zinc. However, the demonstration(s) using simplified equations are easy and most students can understand the basic idea. Below is a scheme which is helpful in teaching about amphiprotic hydroxides:

$$3H^+ + AlO_3^{---} \leftrightharpoons Al(OH)_3 \rightleftharpoons Al^{+++} + 3OH^-$$

Addition of a strong base shifts the equilibrium to the right.

Addition of a strong acid shifts the equilibrium to the left.

The second lesson on bases deals with preparations of hydroxides, basic anhydrides and nomenclature. Use Sections 14.10, 14.11, part of 14.12 and Exp. 24B, "Preparation and Properties of Hydroxides."

Deal with definitions, methods of preparation and nature of salts. Then take up nomenclature and solubility. Emphasize the rules in Sec. 14.15 for acid-salt nomenclature. At this time, some teachers may devote another lesson to salt hydrolysis. Use Exp. 39, "Hydrolysis" and simplify the treatment in Chapter 20.

Chapter 15
ACID-BASE TITRATION AND pH

The practical aspects of this chapter involve:
1. Molar solutions.
2. Titration techniques.
3. pH as a practical scheme for expressing and utilizing acidity of solutions. Wherever possible, students should be trained in the use of pH meters. Experiences should also involve use of a variety of indicators whose color changes are correlated with the use of pH paper.

Topics such as chemical equivalents and normal solutions, as well as, the related mathematics are generally omitted along with all except the simplest pH calculations. Selected portions of experiments which provide the desirable experiences are taken from Exp. 25, "Hydronium Ion Concentration, pH"; Exp. 26, "Brønsted Acids and Bases Indicators"; Exp. 27, "Relative Strengths of Acids and Bases." Exp. 29, "Percentages of Acetic Acid in Vinegar" is one that students enjoy. Do the arithmetic using Molarity instead of Normality.

The slight ionization of water should be demonstrated to establish pH = 7. Students then use the scheme below as the basis for practical use of the pH scale. This may be used in conjunction with a laboratory experiment which selects representative substances from among those listed in the Table, Approximate pH of Some Common Substances.

pH

1 . . . . . 7 14

Acid (neutral) Basic

← increasingly acid

increasingly basic →

UNIT 6
CARBON AND ITS COMPOUNDS

Chapter 16
CARBON AND ITS OXIDES

The details of resonance and hybridization may be omitted. Exp. 31, "Carbon" and Exp. 35, "Carbon Dioxide" are suitable and should be used.

Chapter 17
HYDROCARBONS

Certain sections such as Sec. 17.3, 17.8, 17.11, 17.18, 17.19 may be omitted. Petroleum, petroleum products, and rubber (natural and synthetic) are topics in which students are interested. Use Exp. 33, "Hydrocarbons."

Chapter 18
HYDROCARBON SUBSTITUTION PRODUCTS

Concentrate on substitution and addition reactions. Omit Sec. 18.2. Details of the preparation of particular halides may be omitted. In dealing with alcohols, mention isopropanol. Alcohol reactions should include nonionic nature (litmus paper, conductivity), reactions with HX and esterification. Diethyl ether should be discussed. The production of formaldehyde from methanol and the Fehling's test may be demonstrated. Exp. 37, "Hydrocarbon Substitution Products" includes the last two items as well as esterification experiments.

UNIT 7
CHEMICAL REACTIONS

Chapter 19
REACTION ENERGY AND REACTION KINETICS

Omit most of this chapter. Teachers may wish to include brief, simplified treatment of exothermic and endothermic reaction, stability and heat of formation, the concept of activation energy, and rates of chemical reactions. Teachers may wish to use Exp. 35 "Rate of Chemical Reactions."

Chapter 20
CHEMICAL EQUILIBRIUM

Omit the contents of this chapter with the following exceptions: Sec. 20.1; Sec. 20.2 to acetic acid equation; Sec. 20.4 (apply Le Chatelier's principle to a single example, such as the Haber reaction, so that students can appreciate the practical significance of factor variation and control). Give Sec. 20.5 detailed treatment.

For Sec. 20.9 use simplified equations for weak acid-strong base and weak base-strong acid compounds. Exp. 37, "Equilibrium," and 39 "Hydrolysis," may be used whole or in part.

Chapter 21
OXIDATION-REDUCTION REACTIONS

Although students may be required to learn and be able to use the simple oxidation number rules, the balancing techniques for involved equations should not be required. Sec. 21.1 and 21.2 may be taught. Simple reactions involving composition, decomposition, and single replacement may be used as examples of oxidation-reduction reactions. Focus attention on replacement reactions: metal and acid, metal and salt solutions, halogen replacements and use Table 21-3 as a focal point of the concepts. Electrolysis: electrolytic cells and electroplating may have been taught in relation to Chapter 13, Ionization (see suggestions). Some teachers may wish to teach the elements of electrochemical cells as applied to the dry cell and/or storage battery. Other chapter details probably can be omitted. Exp. 41, "Oxidation-Reduction Reactions," may be used. Parts of Exp. 43B, "Electrolytic Cells," may also be used.

UNIT 8
A TYPICAL PERIOD OF ELEMENTS

Chapter 22
ELEMENTS OF PERIOD THREE

This chapter has great value as an over-all review or summary of periodicity. It is probably strategic to schedule the chapter after representative elements of Groups I—VIII have been studied. In order to control the depth of treatment in this chapter, omit sp^3 hybridization, use formula rather than ionic equations, study equations for preparation of H_2S and HCl, not other hydrides. Use Exp. 44, "Elements of Period Three" whole or parts.

UNIT 9
THE HIGHLY ELECTRO-POSITIVE ELEMENTS

Chapter 23
THE METALS OF GROUP I

Concentrate on generalizations in Sec. 1, important sodium compounds, selected potassium compounds, and elementary aspects of spectroscopy as they apply to flame tests. Teach significant uses of the metals and their compounds.

There may be two lessons on sodium and its compounds. Demonstrate electrolysis of brine and the soda-lime process for NaOH. The Solvay process details are taught as desired. (See the Teacher Notes in Chapter 23 for details.) Schedule Exp. 45, "Flame Tests."

Chapter 24
THE METALS OF GROUP II

Use Exp. 46, "Group II Metals and Their Compounds." Emphasize family relationships and variations. Contrast with Group I metals. Stress the practical aspects of the uses of the metals and their compounds.

UNIT 10
ELEMENTS OF THE CENTRAL REGION

Chapter 25
THE TRANSITION METALS

The concept that the next-to-outermost shell is being built up to 18 electrons for Sc . . . Zn is probably sufficient. Treat the general properties of the transition metals descriptively. Demonstrate where possible. In the first lesson demonstrate the formation of $Cu(NH_3)_4^{++}$ as described in Exp. 40, "Equilibrium and Complex Ions," and/or Exp. 50, "Complex Ions of Copper and Silver." For color change demonstrations use Exp. 49, "Oxidation States of Transition Elements."

The Iron Family needs three lessons. Center the introductory one about the blast furnace, the second one on steel. You may wish to extend the textbook treatment to include some steel alloys and tempering. Demonstrate that a bobby pin is flexible. Heat the pin in the burner flame and plunge into cold water. The pin is readily broken into small pieces. Repeat with a second bobby pin, but permit it to cool slowly in air. This time, flexing permits the pin to be bent into permanent shapes since much of its flexibility has been lost. See the Teacher Notes in Part VII for additional materials on iron and steel.

The experiments which may be used include Exp. 47, "Metallurgy of the Transition Metals," Exp. 48, "Tests for Iron(II) and Iron(III) Ions," and Exp. 52, "Cobalt Nitrate Tests." An entire lesson may be devoted to copper, its compounds, and some mention of its alloys; parts of Exp. 47 are related. The next lesson may be devoted to silver and gold and some treatment of photography may be desirable. The Zinc subgroup may require two lessons. Use Exp. 51, "Zinc and Its Compounds."

Chapter 26
ALUMINUM AND THE METALLOIDS

The subject of aluminum and its compounds may be covered in the first two lessons. Schedule Exp. 54,

"Aluminum and Its Compounds." Two more lessons may be required to complete the chapter. One of these may be devoted to that part of Sec. 27.1 that is applicable, and to the sections on boron with the exception of Sec. 26.12. Schedule Exp. 55, "Borax Bead Tests." The lesson on silicon emphasizes metalloid behavior and compounds including glass. Dependent on class interest and needs this lesson may actually take two periods. Discussion of arsenic and antimony is minimal as required.

UNIT 11
THE HIGHLY ELECTRONEGATIVE ELEMENTS

Chapter 22
NITROGEN AND ITS COMPOUNDS

You may omit Sec. 27.3 part 2. Substitute the production of nitrogen (actually oxygen-free air) by using the demonstration with phosphorus described in Exp. 14, "Percentage of Oxygen in Air." In this demonstration you may test the water for presence of hydronium ion produced from interaction of P_4O_{10} and the water. Some teachers may choose to demonstrate the preparation and collection of NO and NO_2 and the change of NO to NO_2 in air. Nitrogen fixation may be treated as a separate, final lesson under headings such as (a) fixation in nature, and (b) fixation by man. Schedule Exp. 57, "Ammonia, the Ammonium Ion, and the Nitrate Ion."

Chapter 28
SULFUR AND ITS COMPOUNDS

The sections omitted in this chapter are 10, part 3, 12, and portions of 14. Practically all of the rest of the chapter is fundamental. The ionization constants of H_2SO_3 need not be emphasized. Schedule Exp. 58, "Forms of Sulfur," Exp. 59, "Hydrogen Sulfide and Qualitative Analysis" (Procedures 7 and 8 may be omitted), and Exp. 60, "Sulfite and Sulfate Ions." In the lesson on oxides of sulfur (Lesson 3) you may wish to demonstrate reaction of SO_2 on acidified solutions of $KMnO_4$.

Chapter 29
THE HALOGEN FAMILY

In this chapter you may wish to omit the details of dissolving AgCl in ammonia-water solution. The rest of the chapter is the standard treatment. Schedule Exp. 61, "Preparation and Properties of Chlorine," Exp. 62, "Preparation and Properties of Iodine," and Exp. 64, "Halide Ions." Exp. 64, "Identification of Salts," is one which students generally find interesting.

UNIT 12
NUCLEAR REACTIONS

Chapter 30
RADIOACTIVITY

Some chemistry courses may not include this topic. For those which do, practically all the chapter content with the possible exception of Sec. 30.14, The cyclotron and other atom smashers, should be included. Use appropriate parts of Exp. 65, "Radioactivity."

Part III

Answers to Questions and Problems

The questions and problems are labelled as Group A or Group B. Group A are relatively easy. Group B are more difficult.

Answers to Questions

Group A

1. Knowledge of chemistry is essential in other areas of science.
2. (*a*) A solid has definite shape, a liquid does not. (*b*) A liquid has definite volume, a gas does not.
3. Color, odor, solubility, density, hardness, melting point, boiling point.
4. Chemical activity, reactions with air, water, acids, bases.
5. Length—meter, capacity—liter, mass—gram.
6. Milli- 0.001, centi- 0.01, deci- 0.1, deka- 10, hecto- 100, kilo- 1000.
7. Every chemical change involves an energy change.
8. The metric system is a decimal system which provides for simple computations.
9. (*a*) cm^2; (*b*) liter; (*c*) kg; (*d*) mm; (*e*) m/sec or km/sec.
10. The metric system does not lend itself to the convenience of division into halves and quarters.
11. Accuracy pertains to the nearness of a measurement to its accepted value and is expressed in terms of error. Precision pertains to the reproducibility of measurement data and is expressed in terms of deviation.
12. (*a*) Temperature refers to the hotness or coldness of a body, or to the ability of a body to transfer heat to, or acquire heat from, other bodies. Heat is the form in which energy is transferred from one body to another due exclusively to a difference in temperature. (*b*) Temperature is measured in Celsius degrees; heat is measured in calories.

Group B

13. Student's list.
14. (*a*) Student's list. (*b*) Student's list.
15. Both liquids and gases have the characteristic of flowing.
16. The property is chemical if its demonstration involves a change in the identity of the substance. If the identity is not changed, it is physical.
17. Matter occupies space and has inertia, therefore can be felt.
18. (*a*) There are four significant figures. (*b*) The last digit, 4, is a doubtful figure.
19. (*a*) No. (*b*) No. (*c*) Yes. The ratio of mass to volume is useful in distinguishing one material from another. This ratio yields density information.

20. (*a*) 12.65 cm accepted because it contains one digit of doubtful accuracy. (*b*) 12.65 -no unit; 12.6 cm -failed to include one estimated place; 12.655 cm -includes two doubtful digits; 126.55 mm -includes two doubtful digits; 12 cm -two additional significant figures possible.
21. (*a*) 127.50 km; (*b*) 1200 m; (*c*) 90027.00 cm^3; (*d*) 0.00530 g; (*e*) 670 mg; (*f*) 0.0730 g; (*g*) 43.050 liter; (*h*) 300900 kg; (*i*) 0.147 cm; (*j*) 6271.9 cm^2.
22. Both a number and a unit are required to express a physical measurement and confusion and errors can be avoided in the computation by solving for the unit as well as the magnitude of the answer.
23. In the absence of any specific tolerance information we may assume an uncertainty of $\pm 0.1^\circ$.
24. An uncertainty of ± 0.01 g in the precision of the measurement is indicated. This represents the limitation in the reproducibility of measurement data.

Solutions to Problems p. 29

Group A

1. (*a*) 10 mm. (*b*) 10^3 mm. (*c*) 10^6 mm.
2. (*a*) 30.48 cm. (*b*) 200 cm. (*c*) 39.37 in.
3. (*a*) 2000 ml. (*b*) 10^4 ml. (*c*) 1000 liters.
4. (*a*) 4×10^5 mg. (*b*) 4.536×10^5 mg. (*c*) 2000 g.
5. (*a*) Student's answer. (*b*) Student's answer.
6. (*a*) 0.250 liter. (*b*) 250 g. of water.
7. 93,000,000 mi = 9.3×10^7 mi.
8. 0.0000005 cm = 5×10^{-7} cm.

Group B

9. (*a*) 1000 ml. (*b*) 1000 cm^3. (*c*) 10.00 cm. (*d*) 0.1000 m.
10. (*a*) $V = \pi r^2 h = 61.3\ cm^3 = 61.3$ ml. (*b*) 61.3 g of water.
11. 453.6 g – 206.4 g = 247.2 g.
12. h (grad. portion) = V (grad.) $\div \pi r^2$ = 18.9 cm. h (total) = 18.9 + 5.2 = 24.1 cm.
13. (*a*) $D = m/V$
    $m = DV = 13.6$ g/ml $\times$ 8.20 ml = 112 g.
    (*b*) $V = m/D = 120$ g/13.6 g/ml = 8.83 ml.
14. (*a*) 152.20 cm $\times$ 1 m/100 cm = 1.5220 m.
    (*b*) 152.20 cm $\times$ 10 mm/cm = 1522.0 mm.
    (*c*) 152.20 cm $\times$ 1 km/100,000 cm = 0.0015220 km.
    (*d*) 152.20 cm $\times$ 1 in/2.5400 cm = 59.921 in.

15. $$\frac{3.70\ \text{ml} \times 6.02 \times 10^{23}\ \text{mol}}{18.0\ \text{g} \times 1\ \text{ml/g}} = 1.24 \times 10^{23}\ \text{mol}.$$

16. $$\frac{1.24 \times 10^{23}\ \text{mol}}{\frac{1\ \text{mol}}{\text{sec}} \times \frac{8.64 \times 10^4\ \text{sec}}{\text{day}} \times \frac{365\ \text{days}}{\text{yr}}} = \frac{1.24 \times 10^{23}}{3.15 \times 10^7}\ \text{yr} = 3.94 \times 10^{15}\ \text{yr}.$$

Chapter 2

Answers to Questions pp. 44–46

Group A

1. The three general classes of matter are elements, compounds, and mixtures.
2. Matter is anything which occupies space. A substance is a particular kind of matter.
3. A complex substance is composed of two or more elements combined chemically. An elementary substance consists of a single element.
4. (*a*) Elements are classed as metals and nonmetals. (*b*) No, some elements have characteristics intermediate between metals and nonmetals, and are called metalloids.

5. A compound is composed of two or more elements combined chemically in a definite proportion by weight. A mixture consists of two or more substances intermingled in any proportion.
6. Oxygen, silicon, aluminum, iron, and calcium.
7. (*a*) Metals have luster, reflect heat and light, and conduct heat and electricity. They are ductile and malleable. (*b*) Nonmetals are poor conductors of heat and electricity. They may be brittle solids; several are normally gaseous.
8. (*a*) 105 elements are known. (*b*) 92 were known prior to the Atomic Age.
9. The symbol identifies a certain element and stands for one atom of that element.
10. (*a*), (*b*), (*c*) Student answers.
11. In a physical change only physical properties are altered; the identifying properties of the substance remain unchanged. In a chemical change the identifying properties of the substance are altered; a new substance or substances with new properties are formed.
12. By increasing the temperature of the reactants.
13 By the second letter of the symbol.
14. (*a*) The first is always capitalized. (*b*) The second is always lower case.

Group B

15. Difference in solubility. Sand is insoluble in water and sugar is soluble.
16. Place the mixture in water. Separate the sand by filtration and recover the sugar by evaporation of water from the solution of sugar.
17. A solution does not consist of a single substance. The properties of the different substances may be recognized. They may be separated by physical means. The composition can be varied.
18. The weight relationship of constituents is constant.
19. It assures the chemist of the constancy of weight relationship of the elements of any compound.
20. Student answer.
21. Student answer. Separation of the solution by physical means since it is a mixture.
22. In the first case, the solid combined with a gas from the air; in the second, the solid decomposed giving up a gaseous product to the air; in the third, no reaction occurred.
23. Gold, silver, and copper are sufficiently inactive that they do not readily combine with other substances and were thus found free in nature long before processes for recovery of other more active metals were developed.
24. Student's answer, which should recognize that iodine (but not antimony) is soluble in such solvents as carbon tetrachloride and trichloroethane, in which it imparts a characteristic color. By extraction, the iodine is taken up by the solvent. By filtration the antimony is recovered. By evaporation of the solvent, the iodine is recovered.
25. (*a*) Chemical, (*b*) chemical, (*c*) physical, (*d*) chemical, (*e*) physical, (*f*) physical.
26. (*a*), (*b*), and (*d*).
27. Student answers.
28. Evidence of chemical action is the production of heat, light, mechanical energy, or electric energy. The evolution of a gas and the formation of a precipitate are also evidences of chemical reaction.
29. Student answer.
30. Chemical change is indicated by evidence of chemical reaction and evidence of new substances with new properties. Physical change is indicated by change in appearance or other physical properties without loss of identifying properties.
31. Reaction processes are influenced by the tendency to proceed toward a lower energy state and the tendency to proceed toward a more disordered state.
32. The endothermic change (melting) occurs at the higher temperature and is in the direction of higher entropy. The exothermic change (freezing) occurs at the lower temperature and is in the direction of lower energy. It would appear that the influence of the entropy-change factor increases as the temperature of the system rises.

Solutions to Problems p. 46

Group B

1. $D = m/V$, therefore, $V = m/D$

$$V = \frac{4.05 \text{ kg}}{7.87 \text{ g/cm}^3} \times \frac{10^3 \text{ g}}{\text{kg}} = 515 \text{ cm}^3$$

2. $$45\bar{0} \text{ g} \times \frac{\text{kg}}{10^3 \text{ g}} \times \frac{1 \text{ kcal}}{\text{kg C}^\circ} \times 79.0 \text{ C}^\circ = 35.6 \text{ kcal}$$

3. $$\frac{45.5 \text{ kcal}}{1.25 \text{ liters} \times 1 \text{ kg/liter} \times 1 \text{ kcal/kg C}^\circ} = 36.4 \text{ C}^\circ$$

4. $$68 \text{ kcal} \times \frac{1.0 \text{ g}}{2.1 \times 10^{10} \text{ kcal}} = 3.2 \times 10^{-9} \text{ g}$$

5. $$10.0 \text{ g sodium} \times \frac{10\bar{0} \text{ g sodium chloride}}{39.3 \text{ g sodium}} = 25.4 \text{ g sodium chloride}$$

Chapter 3

Answers to Questions pp. 59–60

Group A

1. Even with a microscope one cannot see the particles of sugar in a sugar-water solution. One cannot see particles of gas escaping from a gas valve into the air.
2. The modern atomic theory includes information concerning the structure and properties of atoms, the kinds of compounds they form, and the properties of these compounds. It also includes information about the mass, volume, and energy relationships in reactions between atoms.
3. *1.* All matter is made up of very small particles called atoms. *2.* Atoms of the same element are chemically alike; atoms of different elements are chemically different. *3.* Individual atoms of an element may not all have the same mass. However, the atoms of an element, as it occurs naturally, have, for practical purposes, a definite average mass that is characteristic of the element. *4.* Individual atoms of different elements may have nearly identical masses. However, the atoms of different naturally occurring elements have different average masses. *5.* Atoms are not subdivided in chemical reactions.
4. (*a*) The nucleus and the shells or energy levels. (*b*) Nucleus; protons and neutrons. Energy levels: electrons. (*c*) Proton—positively charged particle with a mass of 1.673×10^{-24} g, which is 1836/1837 of the mass of a protium atom. Neutron—neutral particle with a mass of 1.675×10^{-24} g, which is about the same mass as a proton. Electron—negatively charged particle with a mass of 9.110×10^{-28} g, which is 1/1837 of the mass of a protium atom.
5. The diameter of the nucleus is about 1/10,000 of the diameter of the atom.
6. A region about the nucleus of an atom in which electrons move.
7. Electrons move in paths about the nucleus in such a manner as to effectively occupy the relatively vast empty space around the nucleus.
8. Protium—one proton in nucleus, one electron in first energy level. Deuterium—one proton and one neutron in nucleus, one electron in first energy level. Tritium—one proton and two neutrons in nucleus, one electron in first energy level.
9. (*a*) The number of protons in the nucleus of that atom. (*b*) It indicates the number of electrons in the neutral atom.
10. (*a*) The different varieties of atoms as determined by their nuclear composition. (*b*) Nuclides having the same atomic number; atoms of the same element which have different masses.
11. The mass number is the sum of the number of protons and neutrons.

12. (*a*) 8. (*b*) 17. (*c*) Oxygen-17.
13. Beryllium and fluorine.
14. The mass of an atom expressed relative to the carbon-12 = exactly 12 scale.
15. (*a*) 6.022169×10^{23}. (*b*) The Avogadro number. (*c*) Mole.
16. The average relative mass of the naturally occurring atoms of an element on the carbon-12 = exactly 12 scale.
17. Carbon-12.
18. They indicate relative mass relationships between reacting elements.
19. (*a*) Silver: at. no. 47, at. wt. 107.868. (*b*) Gold: at. no. 79, at. wt. 196.9665. (*c*) Copper: at. no. 29, at. wt. 63.546. (*d*) Sulfur: at. no. 16, at. wt. 32.06. (*e*) Uranium: at. no. 92, at. wt. 238.029.
20. (*a*) 2.00 moles × 4.00260 g/mole = 8.01 g. (*b*) 5.00 moles × 10.81 g/mole = 54.0 g. (*c*) 0.500 mole × 20.179 g/mole = 10.1 g. (*d*) 0.250 mole × 24.305 g/mole = 6.08 g. (*e*) 0.100 mole × 28.086 g/mole = 2.81 g.
21. (*a*) $\frac{20.823\text{ g}}{6.941\text{ g/mole}} = 3.000$ moles. (*b*) $\frac{160.93\text{ g}}{22.9898\text{ g/mole}} = 7.0001$ moles.

    (*c*) $\frac{3.995\text{ g}}{39.948\text{ g/mole}} = 0.1000$ mole. (*d*) $\frac{8.016\text{ g}}{32.06\text{ g/mole}} = 0.2500$ mole.

    (*e*) $\frac{20.24\text{ g}}{26.9815\text{ g/mole}} = 0.7501$ mole.

22. The law of definite composition; and the way and the proportions in which substances react with one another.
23. Most of a platinum atom is space through which high-speed positively charged particles readily pass. However, platinum atoms contain a very small, positively charged nucleus which repels the high-speed positively charged particles if they approach head-on or very nearly so.
24. (*a*) The atoms of each successive element have one more proton. (*b*) The atoms of each successive element have one more electron. (*c*) There is no definite pattern, but generally the number of neutrons increases.
25. Two electrons in the first energy level and from one to eight electrons in the second energy level.
26.

Name of nuclide	Atomic number	Mass number	Composition of nucleus		Electron configuration		
			Protons	Neutrons	K	L	M
sodium-23	11	23	11	12	2	8	1
magnesium-24	12	24	12	12	2	8	2
aluminum-27	13	27	13	14	2	8	3
silicon-28	14	28	14	14	2	8	4
phosphorus-31	15	31	15	16	2	8	5
sulfur-32	16	32	16	16	2	8	6
chlorine-35	17	35	17	18	2	8	7
argon-40	18	40	18	22	2	8	8

27. Approximately 75% chlorine-35 and 25% chlorine-37.
28. They will be the same.
29. (*a*) They are isotopes. (*b*) They are of different elements but have the same mass number.

30. Divide the atomic mass in grams by the Avogadro number.
31. It is the weighted average mass of an atom of the element in grams.

32. $0.99759 \text{ mole} \times \frac{15.99491 \text{ g}}{1 \text{ mole}} = 15.9564 \text{ g}$

$0.00037 \text{ mole} \times \frac{16.99914 \text{ g}}{1 \text{ mole}} = 0.0063 \text{ g}$

$0.00204 \text{ mole} \times \frac{17.99916 \text{ g}}{1 \text{ mole}} = 0.0367 \text{ g}$

15.9564 g + 0.0063 g + 0.0367 g = 15.9994 g, the gram-atomic weight. The atomic weight is 15.9994.

Chapter 4

Answers to Questions pp. 72–74

Group A

1. (*a*) Electromagnetic radiations are forms of energy which travel through space as waves at the rate of 3.00×10^8 m/sec. (*b*) Light, X rays, radio waves.
2. (*a*) They travel through space as waves. (*b*) They are transferred to matter in quanta (units) of energy called photons.
3. By the absorption of a photon.
4. (*a*) 3rd energy level to 2nd energy level. (*b*) 4th energy level to 2nd energy level.
5. A nucleus surrounded by electrons moving in definite energy levels—that is, only at certain distances from the nucleus and with only certain speeds.
6. A highly probable location about a nucleus in which an electron may be found.
7. (*a*) The space surrounding a nucleus bounded by the occupied space orbitals. (*b*) It gives size and shape to an atom and prevents two free atoms, or portions of free atoms, from occupying the same space.
8. The principal quantum number indicates the average distance of the electron from the nucleus of the atom. The orbital quantum number indicates the shape of the space orbital. The magnetic quantum number indicates the position in space of the space orbital. The spin quantum number indicates the two possibilities for the property of an electron called electron spin.
9. (*a*) A hollow sphere. (*b*) One. (*c*) Two. (*d*) They must have opposite spin. (*e*) The first.
10. (*a*) Approximately that of two footballs or watermelons each tangent to the other at an end; two ellipsoids each tangent to the other at a point of greatest curvature. (*b*) Three. (*c*) At 90° angles along the three axes in space. (*d*) The second.
11. (*a*) No. (*b*) Yes. (*c*) Their first three quantum numbers must be alike, but they must have opposite spins.
12. In its ground state, the electrons of an atom occupy the lowest energy orbitals available. In an excited state, one or more electrons of an atom occupy orbitals of higher energy.
13. (*a*) Two electrons of opposite spin occupying the same space orbital. (*b*) Eight electrons occupying the *s* and *p* orbitals of an outer energy level.

Group B

14. The attraction between oppositely charged particles requires that atoms collapse. Since this is not so, this explanation is unsatisfactory.
15. $c = f\lambda$. So $f = c/\lambda$ $E = hf$. So $f = E/h$, Then $c/\lambda = E/h$. And $\lambda = ch/E$.
16. Because electrons may only possess definite amounts of energy.
17. (*a*) 1000 Å. (*b*) 18,800 Å.

18. (*a*) Five. (*b*) Ten. (*c*) The third.
19. (*a*) Seven. (*b*) Fourteen. (*c*) The fourth.
20.

Chemical symbol	Orbital notation					Electron-configuration notation	Electron-dot notation
	1*s*	2*s*	2*p*	3*s*	3*p*		
Na	⊗	⊗	⊗ ⊗ ⊗	⊘	○ ○ ○	$1s^2 2s^2 2p^6 3s^1$	Na·
Mg	⊗	⊗	⊗ ⊗ ⊗	⊗	○ ○ ○	$1s^2 2s^2 2p^6 3s^2$	Mg:
Al	⊗	⊗	⊗ ⊗ ⊗	⊗	⊘ ○ ○	$1s^2 2s^2 2p^6 3s^2 3p^1$	Al:
Si	⊗	⊗	⊗ ⊗ ⊗	⊗	⊘ ⊘ ○	$1s^2 2s^2 2p^6 3s^2 3p^2$	·Si:
P	⊗	⊗	⊗ ⊗ ⊗	⊗	⊘ ⊘ ⊘	$1s^2 2s^2 2p^6 3s^2 3p^3$	·P:
S	⊗	⊗	⊗ ⊗ ⊗	⊗	⊗ ⊘ ⊘	$1s^2 2s^2 2p^6 3s^2 3p^4$	·S:
Cl	⊗	⊗	⊗ ⊗ ⊗	⊗	⊗ ⊗ ⊘	$1s^2 2s^2 2p^6 3s^2 3p^5$	:Cl:
Ar	⊗	⊗	⊗ ⊗ ⊗	⊗	⊗ ⊗ ⊗	$1s^2 2s^2 2p^6 3s^2 3p^6$	:Ar:

21. (*a*) One. (*b*) Four. (*c*) Two. (*d*) One. (*e*) Three.
22. Krypton.
23. Seven.
24. The completion of *s* and *p* sublevels of the outer energy levels is interrupted by the occupation of the *d* sublevel of the next-to-outside energy level.
25. The completion of *s* and *p* sublevels of the outer energy level and of the *d* sublevel of the next-to-outside energy level is interrupted by the occupation of the *f* sublevel in the second-from-outside energy level.
26. (*a*) Fourth. (*b*) *s*, *p*, *d*, and *f*. (*c*) One *s*, three *p*'s, five *d*'s, and seven *f*'s. (*d*) Two in the *s* orbital, six in the *p* orbitals, ten in the *d* orbitals, and fourteen in the *f* orbitals. (*e*) 32.
27. (*a*) *s* and *p*. (*b*) *s*, *p*, and *d*.
28. $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10} 4f^{14} 5s^2 5p^6 5d^{10} 5f^{14} 6s^2 6p^6 6d^4 7s^2$.

Chapter 5
Answers to Questions pp. 89–90

Group A

1. (*a*) In order of increasing atomic weights. (*b*) In order of increasing atomic numbers.
2. He predicted the discovery of new elements and their properties.
3. The wavelengths of X rays become shorter as the elements used as a target contain more protons in their nuclei.
4. The physical and chemical properties of the elements are periodic functions of their atomic numbers.
5. (*a*) Chemical symbol, atomic number, atomic weight, distribution of electrons in the energy levels of the atoms of the element. (*b*) The chemical symbol is in the center with the atomic number below, the atomic weight above, and the electron distribution to the right.
6. (*a*) A vertical column of elements in the periodic table. (*b*) One of the vertical columns.
7. (*a*) A horizontal row of elements in the periodic table. (*b*) One of the horizontal rows.

8. (*a*) Lithium, beryllium, boron, carbon, nitrogen, oxygen, fluorine, neon. (*b*) The number increases from one to eight as you go across the series from lithium to neon. (*c*) The properties range from an active metallic element to an active nonmetallic element, with the last element in the period being a noble gas.
9. They have the same number of electrons in their outer energy levels.
10. They increase in activity from the top to the bottom of a group.
11. They decrease in activity from the top to the bottom of the group.
12. Metalloids.
13. The boundary of an atom's electron cloud is not a distinct surface because the space occupied by an atom is that taken by charged electrons moving rapidly about a nucleus.
14. (*a*) The radii of the Group I elements are greatest among the elements of their period. (*b*) The single outer energy level electron is rather loosely bound to the nucleus.
15. (*a*) $K + \text{energy} \rightarrow K^+ + e^-$. (*b*) $Br + e^- \rightarrow Br^- + \text{energy}$. (*c*) A singly positively charged potassium ion and a singly negatively charged bromide ion.
16. (*a*) The few outer shell electrons of the metal atoms are not strongly held by nuclear attraction. (*b*) The outer shell electrons of nonmetal atoms are more strongly held by nuclear attraction.

Group B

17. The noble gases.
18. (*a*) Electromagnetic radiations which have a high frequency and a short wavelength. (*b*) By high-speed electrons striking the metal target in an evacuated tube.
19. (*a*) By studying the properties of the elements he already knew. (*b*) By unusual variations in the wavelengths of X rays between two supposedly successive elements.
20. (*a*) Because of its many unique properties. (*b*) Because it has one electron in its outermost energy level.
21. (*a*) Elements which usually differ in electronic configuration from that of the next lower or higher atomic number only in the number of electrons in the next-to-the-outside energy level. (*b*) Fourth, fifth, sixth, and seventh periods.
22. (*a*) Elements which usually differ in electronic configuration from that of the next lower or higher atomic number only in the number of electrons in the second-from outside energy level. (*b*) Sixth and seventh periods.
23. (*a*) Atomic size increases with atomic number within a family of elements. (*b*) Each element in a family has one more energy level than the element above it. The addition of an energy level increases atomic size more than increased nuclear charge reduces it.
24. (*a*) Atomic size generally decreases with atomic number. (*b*) The attraction of the increased nuclear charge for the electrons pulls them closer to the nucleus.
25. (*a*) The atom with higher atomic number should have the higher ionization energy. (*b*) The nuclear attraction for its outer energy level electrons should be greater.
26. (*a*) The atom is more stable because it has less energy than the ion. (*b*) The ion is more stable because it has less energy than the atom.
27. The number of orbitals which have energies between the completed *p* sublevels of successive noble gas configurations. The number of elements in a period is twice the number of such orbitals.
28.

Chemical element	Electron configuration notation
K	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$
Ca	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$
Ga	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^1$

In general, the ionization energy increases with each electron removed from an atom,

because each successive electron must be removed from a particle with an increasingly greater net positive charge. Also, the lower the energy level of an electron, the more ionization energy is required for its removal. Potassium: The first electron is a 4*s* electron, in a higher energy level than the second and third electrons, which are 3*p* electrons. Calcium: The first two electrons are 4*s* electrons, in a higher energy level than the third and fourth electrons, which are 3*p* electrons. Gallium: The first electron is a 4*p* electron. The second and third are 4*s* electrons, in a lower energy sublevel than the first. The fourth electron is a 3*d* electron, in a still lower energy level.

29. 16.
30. It serves as a useful and systematic, though not perfect, classification of elements according to their properties.

Chapter 6

Answers to Questions pp. 117–119

Group A

1. (*a*) The electrons in an incomplete outer shell; the valence electrons. (*b*) Usually by the transfer of valence electrons from the outer shell of one atom to the outer shell of another atom, or by the sharing of valence electrons among the outer shells of combining atoms.
2. (*a*) Ionic bonding and covalent bonding. (*b*) Ionic bonding produces ions; covalent bonding produces molecules.
3. (*a*) A stable one consisting of an octet. (*b*) Because atoms with such outer shells have no tendency to transfer and little tendency to share electrons.
4. (*a*) Exothermic. (*b*) Usually exothermic, but may be endothermic.
5. (*a*) A chemical formula is a shorthand method of using chemical symbols to represent the composition of a substance. (*b*) An empirical formula merely indicates the kind and simplest whole number ratio of the constituent atoms in a substance. A molecular formula indicates the actual composition of a molecule.
6. (*a*) An empirical formula. (*b*) A molecular formula.
7. (*a*) $K\cdot$ (*b*) K^+
8. (*a*) $\cdot\overset{\cdot\cdot}{\underset{\cdot\cdot}{S}}\cdot$ (*b*) $:\overset{\cdot\cdot}{\underset{\cdot\cdot}{S}}:^{--}$
9. K^+ $:\overset{\cdot\cdot}{\underset{\cdot\cdot}{S}}:$ $^{--}$ K^+
10. (*a*) Any chemical reaction which involves the loss of one or more electrons by an atom or ion. (*b*) Any chemical reaction which involves the gain of one or more electrons by an atom or ion.
11. (*a*) Because it transfers electrons to the substance being reduced. (*b*) Because it receives electrons from the substance being oxidized.
12. A barium atom has electrons in two or more energy levels than a calcium atom, and the effect of additional energy levels more than offsets the greater nuclear charge of the barium atom.
13. A bromide ion has electrons in one more energy level than a chloride ion. The effect of the addition of an energy level is greater than the effect of greater nuclear charge.
14. (*a*)

	1*s*	2*s*	2*p*
F	⊗	⊗	⊗ ⊗ ⊘
F	⊗	⊗	⊗ ⊗ ⊘

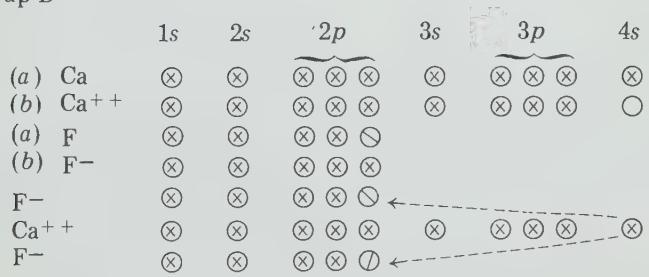
(*b*) $:\overset{\cdot\cdot}{\underset{\cdot\cdot}{F}}:$ $\overset{\cdot\cdot}{\underset{\cdot\cdot}{F}}$

(*c*) F_2

15. An atom is the smallest particle of an element that can exist either alone or in combination with other atoms like it or different from it. A molecule is the smallest chemical unit of an element or a compound which is capable of stable independent existence.
16. A symbol represents one atom of an element, while a formula represents the composition of a substance.
17. (*a*) N_2 is most stable. (*b*) Cl_2 is least stable.
18. (*a*) Hydrogen, nitrogen, oxygen, fluorine, chlorine. (*b*) H_2, N_2, O_2, F_2, Cl_2.
19. (*a*) Helium, neon, argon, krypton, xenon, radon. (*b*) He, Ne, Ar, Kr, Xe, Rn.
20. Sugar: 12 atoms of carbon, 22 atoms of hydrogen, 11 atoms of oxygen. Sand: 1 atom of silicon, 2 atoms of oxygen. Salt: 1 atom of sodium, 1 atom of chlorine. Hydrogen peroxide: 2 atoms of hydrogen, 2 atoms of oxygen. Soap: 18 atoms of carbon, 35 atoms of hydrogen, 2 atoms of oxygen, 1 atom of sodium.
21. (*a*) 1 atom of argon. (*b*) 4 molecules of nitrogen, each molecule of which consists of two atoms of nitrogen. (*c*) 1 molecule of a substance, each molecule of which consists of 1 atom of hydrogen and 1 atom of iodine. (*d*) 6 molecules of a substance, each molecule of which consists of 2 atoms of hydrogen, 1 atom of sulfur, and 4 atoms of oxygen. (*e*) 3 atoms of copper. (*f*) 2 chemical units, each consisting of 1 potassium ion and 1 bromide ion. (*g*) 1 molecule of a substance, each molecule of which consists of 1 atom of carbon and 1 atom of oxygen. (*h*) 1 atom of cobalt.
22. The property of an atom attracting the shared electrons forming a bond between it and another atom.
23. (*a*) Approximately 1.7 or more. (*b*) Greater than about 0.5 but less than about 1.7. (*c*) Less than about 0.5.
24. (*a*) Sodium. (*b*) Chlorine. (*c*) Chlorine. (*d*) Lithium. (*e*) Hydrogen.
25. (*a*) Mn +4, O −2. (*b*) H +1, P +5, O −2. (*c*) H +1, N +5, O −2. (*d*) P +5, O −2. (*e*) Na +1, O −2, H +1.

Group B

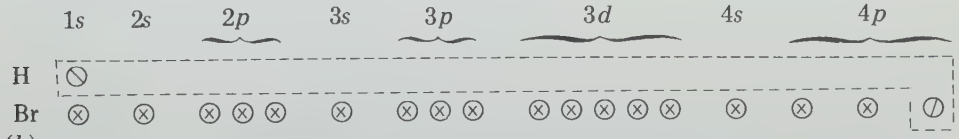
26.



27.

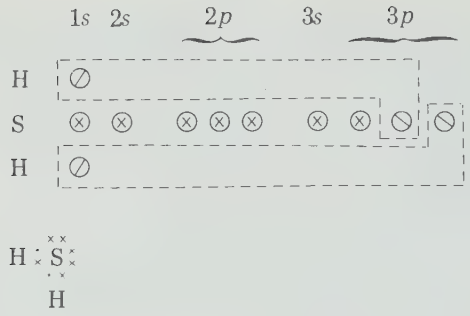
28.

29. The greater nuclear charge of the calcium atom draws the 4*s* electrons closer to the nucleus.
30. Each has an outer shell consisting of an octet of M-shell electrons. The nuclear charge of the sulfur is weaker, thus the octet of this ion is further from the nucleus.
31. (*a*)



(*b*)

32.



33. (*a*) (*b*) (*c*) It must be a resonance hybrid.

34. (*a*) Pyramidal, similar to NH_3. (*b*) Bent, similar to H_2O.
35. (*a*) 63%, ionic. (*b*) 22%, polar covalent. (*c*) 82%, ionic. (*d*) 4%, essentially pure covalent. (*e*) 0%, pure covalent.
36. (*a*) Ionic crystal. (*b*) Nonpolar covalent molecule. (*c*) Polar covalent molecule. (*d*) Nonpolar covalent molecule. (*e*) Nonpolar covalent molecule.
37. (*a*) +7. (*b*) +2. (*c*) Reduction. (*d*) Oxidizing agent.
38. $HClO$, +1; $HClO_2$, +3; $HClO_3$, +5; $HClO_4$, +7.
39. (*a*) S +6, O −2. (*b*) S +4, O −2. (*c*) N +3, O −2. (*d*) C +4, O −2. (*e*) Cr +6, O −2.

Solutions to Problems pp. 119–120

Group A

1. $\frac{1}{2}$ mole H_2 + 52 kcal → 1 mole H
   $\frac{1}{2}$ mole Br_2 + 23 kcal → 1 mole Br
   1 mole H + 1 mole Br → 1 mole HBr + 87 kcal
   $\frac{1}{2}$ mole H_2 + $\frac{1}{2}$ mole Br_2 → 1 mole HBr + 12 kcal
   Energy change is exothermic; 12 kcal/mole HBr.
   (Note to teacher: The reason this heat of reaction does not equal the heat of formation of hydrogen bromide is because we have not considered the approximately 3 kcal required to convert $\frac{1}{2}$ mole Br_2 (l) to $\frac{1}{2}$ mole Br_2(g). In the elementary bond energy calculations in this chapter all substances are in the gaseous state for the sake of simplicity.)
2. 1 mole H_2 + 104 kcal → 2 moles H
   $\frac{1}{2}$ mole O_2 + 60 kcal → 1 mole O
   2 moles H + 1 mole O → 1 mole H_2O + 222 kcal
   1 mole H_2 + $\frac{1}{2}$ mole O_2 → 1 mole H_2O + 58 kcal
   Exothermic energy change; 58 kcal/mole H_2O.

Group B

3. 1 mole Mg + 521 kcal → 1 mole Mg^{++} + 2 moles e^-
   2 moles Br + 2 moles e^- → 2 moles Br^- + 165 kcal
   1 mole Mg^{++} + 2 moles Br^- → 1 mole $Mg^{++}Br_2^-$ + 575 kcal
   1 mole Mg + 2 moles Br → 1 mole Mg^{++} Br_2^- + 219 kcal
   Exothermic energy change; 219 kcal/mole $MgBr_2$.

4. $1 \text{ mole } H_2 + 104 \text{ kcal} \rightarrow 2 \text{ moles H}$
$1 \text{ mole } O_2 + 119 \text{ kcal} \rightarrow 2 \text{ moles O}$
$2 \text{ moles H} + 2 \text{ moles O} \rightarrow 2 \text{ moles H—O} + 222 \text{ kcal}$
$2 \text{ moles H—O} \rightarrow 1 \text{ mole } H_2O_2 + 34 \text{ kcal}$

$1 \text{ mole } H_2 + 1 \text{ mole } O_2 \rightarrow 1 \text{ mole } H_2O_2 + 33 \text{ kcal}$

Exothermic energy change; 33 kcal/mole H_2O_2.

5. $1 \text{ mole } N_2 + 227 \text{ kcal} \rightarrow 2 \text{ moles N}$
$2 \text{ moles } H_2 + 208 \text{ kcal} \rightarrow 4 \text{ moles H}$

$2 \text{ moles N} + 4 \text{ moles H} \rightarrow 2 \text{ moles } \mathrm{N}\begin{smallmatrix}\diagup\mathrm{H}\\ \diagdown\mathrm{H}\end{smallmatrix} + 4(92 \text{ kcal})$

$2 \text{ moles } \mathrm{N}\begin{smallmatrix}\diagup\mathrm{H}\\ \diagdown\mathrm{H}\end{smallmatrix} \rightarrow 1 \text{ mole } N_2H_4 + 37 \text{ kcal}$

$1 \text{ mole } N_2 + 2 \text{ moles } H_2 + 30 \text{ kcal} \rightarrow 1 \text{ mole } N_2H_4$

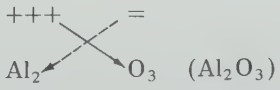
Energy change is endothermic; 30 kcal/mole N_2H_4.
(Note to teacher: In this problem the heat evolved when $N_2H_4(g)$ condenses to $N_2H_4(l)$ has not been considered. The heat of formation of $N_2H_4(l)$ is +12 kcal/mole.)

Chapter 7

These exercise sheets referred to in the Teacher's Notes are intended for use with Sections 7.2 and 7.3.

Formula Writing
Exercise Sheet 1

1. Ionic formulas.
(a) The combining capacity of an ion depends on the magnitude of its electrical charge, that is, the *number* of positive or negative *charges* the ion has.
(b) The charge on an ion may be determined from the electron configuration of its atom. For monatomic ions, the electrical charge corresponds to one of the *oxidation states* listed in the Periodic Table. (See Table for Quiz 1 and Quiz 2 on following page.)
(c) The principle to be used for writing formulas from known charges is called the principle of equality of total positive and total negative charge. This principle evolved and was apparent whenever you constructed models such as atom models for ionic substances and wrote the resultant formula.
(d) Rules for formula writing.
(1) Place the symbols for the ions side by side, e.g. Na Cl.
(2) Place the ionic charge above each symbol.
(3) If the number of positive charges is equal to the number of negative charges, the previous procedure gives the correct formula, e.g. NaCl, MgO.
(4) If the number of positive and negative charges is *not identical*, then *cross-over the number* of each kind of charge, using this number as a subscript (In order to achieve equality of charge)



(5) If a radical is to be taken more than once, it must be enclosed in parenthesis, i.e., calcium hydroxide: $Ca(OH)_2$.
The Exercise Sheet #2 may be used at this point, if desired, to drill concepts of oxidation number (state) as they relate to such formulas of molecular substances as HCl,

HBr, NH_3, CH_4, CO_2, and CCl_4. This also may be used in conjunction with Section 4. Naming compounds from their formulas, as well as Section 7.5, Significance of chemical formulas (empirical, molecular). Redrill the use of coefficients with formulas in relation to Section 7.6, Molecular weight and Section 7.7, Formula weight of a compound.

Quiz 1

	Symbol or Radical	Number and Kind-charge (Oxidation No)
1. aluminum	.....	.....
2. ammonium	.....	.....
3. antimony	.....	.....
4. barium	.....	.....
5. calcium	.....	.....
6. chloride	.....	.....
7. copper(II)	.....	.....
8. iron(II)	.....	.....
9. hydroxide	.....	.....
10. mercury(I)	.....	.....
11. carbonate	.....	.....
12. magnesium	.....	.....
13. lead	.....	.....
14. mercuric	.....	.....
15. iron(III)	.....	.....
16. oxide	.....	.....
17. nitrate	.....	.....
18. sulfide	.....	.....
19. phosphate	.....	.....
20. sodium	.....	.....
21. potassium	.....	.....
22. zinc	.....	.....
23. silver	.....	.....
24. sulfate	.....	.....
25. hydronium	.....	.....
26. hydrogen carbonate	.....	.....
27. bromide	.....	.....
28. chlorate	.....	.....
29. sulfite	.....	.....
30. fluoride	.....	.....
31. hypochlorite	.....	.....
32. hydride	.....	.....
33. iodide	.....	.....

Quiz 2

	Formulas
1. sodium chloride	.....
2. ammonium hydroxide	.....
3. calcium sulfate	.....
4. magnesium nitrate	.....
5. aluminum phosphate	.....
6. zinc chloride	.....
7. mercury(II) oxide	.....
8. aluminum sulfate	.....
9. silver nitrate	.....
10. barium hydroxide	.....
11. potassium sulfide	.....
12. iron(II) sulfate	.....
13. mercury(I) chloride	.....
14. copper carbonate	.....
15. aluminum phosphate	.....
16. iron(III) sulfate	.....
17. calcium phosphate	.....
18. zinc sulfide	.....
19. ammonium carbonate	.....
20. antimony chloride	.....
21. potassium oxide	.....
22. ammonium sulfide	.....
23. mercuric nitrate	.....
24. iron(III) chloride	.....
25. aluminum oxide	.....

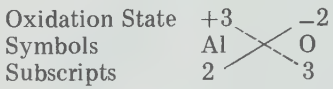
NOTE: #26-33 may be used as part of second Quiz on Ions and Charges

	Negative ions (including Radicals)							
	chloride (Cl)	nitrate (NO_3)	hydroxide (OH)	oxide (O)	sulfide (S)	sulfate (SO_4)	carbonate (CO_3)	phosphate (PO_4)
Positive ions								
sodium (Na)	NaCl	$NaNO_3$	NaOH	Na_2O	Na_2S	Na_2SO_4	Na_2CO_3	Na_3PO_4
potassium (K)	KCl							
silver (Ag)	AgCl							
ammonium (NH_4)	NH_4Cl							
mercury(I) (Hg_2) (ous)	Hg_2Cl_2							
mercury(II) (Hg) (ic)	$HgCl_2$							
calcium (Ca)	$CaCl_2$							
magnesium (Mg)	$MgCl_2$							
barium (Ba)	$BaCl_2$							
zinc (Zn)	$ZnCl_2$							
copper(II) (Cu) (ic)	$CuCl_2$							
lead(II) (Pb)	$PbCl_2$							
iron(II) (Fe) (ous)	$FeCl_2$							
iron(III) (Fe) (ic)	$FeCl_3$							
aluminum (Al)	$AlCl_3$							
antimony(III) (Sb)	$SbCl_3$							

Oxidation States and Formulas

Exercise Sheet 2

1. Oxidation in the simples' sense means chemical combination with oxygen.
   $2Mg + O_2 \rightarrow 2MgO$ $2H_2 + O_2 \rightarrow 2H_2O$
   $4Al + 3O_2 \rightarrow 2Al_2O_3$ $C + O_2 \rightarrow CO_2$
2. Examination of the electron diagrams of the atoms involved in the formation of ionic compounds such as MgO and Al_2O_3 reveals that the metals involved have lost one or more electrons during the oxidation:
   $Mg - 2e \rightarrow Mg^{++}$ $Al - 3e \rightarrow Al^{+++}$
   - 2.1 When a metal such as Mg or Na burns in chlorine, a similar loss of electrons takes place.
   - 2.2 Therefore, a more comprehensive definition of oxidation is: *oxidation* is the *loss* of one or more electrons.
   - 2.3 The metals, therefore exist in characteristic *oxidation states* in ionic compounds. The *oxidation state* is the signed number of charges on the ion. Oxidation states of metals: Al^{+++} or +3; Mg^{++} or +2; Na^{+} or +1.
   - 2.4 During oxidation resulting in ionic compounds, the nonmetals gain one or more electrons (this process is sometimes called reduction). They consequently develop one or more negative charges and therefore have negative oxidation states.
     $O + 2e^- \rightarrow O^=$, Oxidation state = −2
     $Cl + e^- \rightarrow Cl^-$, Oxidation state = −1
   - 2.5 In writing formulas of Ionic Compounds we used these oxidation states to arrive at the correct formulas.



   We used the *principle of equality of charge* to arrive at the formula.
3. In writing formulas for nonionic, that is molecular substances, the same principle may be employed by using a system of assigning oxidation states to the elements (uncombined and/or in compounds.
   - 3.1 This assignment of oxidation states is logical when it is evident that both carbon and hydrogen combine with oxygen, i.e., are oxidized when they burn to form carbon dioxide and water respectively.
   - 3.2 Oxidation state rules.
     (a) The Oxidation state of an atom of a free element is zero.
     (b) The oxidation state of monatomic ions is equal to their ionic charge.
     (c) In combinations between nonmetals, the more electronegative element is assigned an oxidation state which is negative. The other element therefore has a positive oxidation state.
     (d) The oxidation state of hydrogen is +1. (In metallic hydrides it is −1)
     (e) The oxidation state of oxygen is almost always −2. The exceptions are
     1. combined with F as in OF_2, the O is assigned a +2.
     2. in peroxides (or peroxide ions) the oxidation state is −1 (H_2O_2).
     (f) The algebraic sum of the oxidation states of a particular group of atoms, whether they be the same or different atoms is termed the *oxidation number.*
     (g) In the formula of a compound, the algebraic sum of the oxidation numbers of the atoms is zero.
   - 3.3 Determining Oxidation States from *given formulas* of compounds.

Oxidation state:	+1	?	?	−2	?	−2	?	+1	?	+1
Symbols:	H_2	S	S	O_2	S	O_3	C	H_4	N	H_3
Oxidation number:	(+2)+?=0		?+(−4)=0		?+(−6)=0		?+(+4)=0		?+(+3)=0	

   - 3.4 Determining Oxidation States in polyatomic ions.
     (a) The algebraic sum of the oxidation numbers must equal the charge on the ion.

Oxidation state	?	−2	For SO_4 =	?	−2	For NO_3^-	?	−2	CO_3 =
Symbols	S	O_4		N	O_3		C	O_3	
Oxidation number?	+ (−8) = −2 (ionic charge)			? + (−6) = −1			? + (−6) = −2		

Example: The oxidation state of P in the phosphate, PO_4^{-3} ion is +5. What is the oxidation number of the PO_4^{-3} ion?

Oxidation state	+5	−2 Ion charge
Symbols	P	O_4

Oxidation numbers (+5) + (−8) = −3

(b) Conversely, the charge on a polyatomic ion is its oxidation number. This oxidation number can be calculated from the oxidation states of the elements composing the ion.

Oxidation state	+1	?	−2	+1	?	−2
Symbols	H_2	S	O_4	H_2	S	O_3
Oxidation number	(+2)+?+(−8)=0			(+2)+?+(−6)=0		

3.5 To determine the Oxidation State of an element in a formula containing different kinds of atoms: Generally such formulas contain O in the typical −2 state. A second element may be hydrogen with its usual oxidation state of +1 or metals of Group IA: oxidation state +1; or of Group IIA: oxidation state +2. The same procedures are followed as for 3.3 (above).

3.6 EXERCISES: Determine the oxidation state of the underlined elements.

(a) $\underline{C}O_2$; $\underline{C}H_4$; $\underline{C}_2H_6$; $\underline{C}_2H_2$; $H_2\underline{C}O_3$; $\underline{C}O_3^=$; $Ca\underline{C}O_3$; $\underline{C}H_2O$; $\underline{C}Cl_4$.
(b) $\underline{N}H_3$; $\underline{N}_2H_4$; $\underline{N}H_4^+$; $\underline{N}H_4OH$; $\underline{N}_2O$; $\underline{N}O$; $\underline{N}O_2$; $\underline{N}O_3^-$; $H\underline{N}O_3$.
(c) $\underline{Mn}Cl_2$; $\underline{Mn}_2O_3$; $\underline{Mn}O_2$; $\underline{Mn}O_4^-$; $K\underline{Mn}O_4$.
(d) $\underline{P}H_3$; $\underline{P}O_4^{-3}$; $H_3\underline{P}O_4$; $NaH_2\underline{P}O_4$; $H_3\underline{P}O_3$.
(e) $H_2\underline{O}$; $Na\underline{O}H$; $H_2\underline{O}_2$; $Na_2\underline{O}_2$; $\underline{O}F_2$; $\underline{O}_2H^-$.
(f) $\underline{Cr}_2O_3$; $\underline{Cr}O_3$; $\underline{Cr}_2O_7^=$; $K_2\underline{Cr}_2O_7$.
(g) $H\underline{Cl}$; $H\underline{Cl}O$; $\underline{Cl}O_2$; $H\underline{Cl}O_2$; $H\underline{Cl}O_3$; $\underline{Cl}O_4^-$; $Na\underline{Cl}O$; $K\underline{Cl}O_3$.

Another lesson may be devoted to Sections 7.6-8, dealing with calculations of molecular and formula weights and with related problem type: Percentage composition of a compound.

This exercise is intended for use with Section 7.11.

(a)

Mass	*No. of moles of Atoms*	*No. of atoms*
130 g Zn		
32.5 g Zn		

(b)

Mass	*Moles of Molecules*	*No. of Molecules*	*No. of H atoms*	*No. of Other Atoms*
18 g H_2O				
36 g H_2O				
9 g H_2O				
4 g CH_4				

(c)

Mass	*Moles of Formula Units*	*No. of Formula Units*	*No. of Cations*	*No. of Anions*
58 g NaCl				
55 g $CaCl_2$				
204 g Al_2O_3				

(d) Calculate the mass in grams represented by the given number of moles

1. 1.0 mole oxygen atoms
2. 0.5 mole H_2SO_4
3. 1.0 mole Al_2O_3
4. 1.0 mole Na^+ ions
5. 1.0 mole oxygen molecules
6. 2.0 moles H_2SO_4
7. 0.25 mole $CuSO_4$
8. 0.5 mole Cl^- ions

Summarize by asking for the two quantities represented by a symbol, a formula for an element (diatomic gases and O_3, P_4, S_8), and a formula for a compound.

Chapter 7

Answers to Questions p. 140

Group A

1. The molecular formula, NH_3, represents the composition of one molecule of ammonia—one atom of nitrogen and three atoms of hydrogen. Since the atomic weight of nitrogen is 14.0 and that of hydrogen is 1.0, the molecular weight of ammonia is 17.0. The mass of one mole of ammonia is 17.0 g.
2. Because it applies to all types of substances, not just to those which are molecular.
3. (*a*) Na^+; (*b*) Cu^+; (*c*) Fe^{+++}; (*d*) Ni^{++} ; (*e*) Pb^{++}.
4. (*a*) Ammonium, +1; (*b*) sulfate, –2; (*c*) nitrate, –1; (*d*) carbonate, –2; (*e*) acetate, –1.
5. (*a*) HCO_3^-; (*b*) Br^-; (*c*) CrO_4^{--}; (*d*) SO_3^{--}; (*e*) PO_4^{---}
6. (*a*) $BaCl_2$; (*b*) CaO; (*c*) $MgSO_4$; (*d*) AgBr; (*e*) $ZnCO_3$.
7. (*a*) Sodium hydrogen carbonate; (*b*) potassium peroxide; (*c*) mercury(II) chloride; (*d*) iron(III) hydroxide; (*e*) nickel(II) acetate.
8. (*a*) NH_4NO_3; (*b*) Al_2S_3; (*c*) $Cu(OH)_2$; (*d*) $Pb_3(PO_4)_2$; (*e*) $Fe_2(SO_4)_3$.
9. (*a*) Copper(II) chloride; (*b*) calcium sulfide; (*c*) potassium hydrogen sulfate; (*d*) sodium nitrite; (*e*) nickel(II) phosphate.
10. (*a*) CrF_3; (*b*) $Ni(ClO_3)_2$; (*c*) $KHCO_3$; (*d*) $CaCrO_4$; (*e*) HgI_2.
11. (*a*) Sodium peroxide; (*b*) ammonium nitrite; (*c*) magnesium phosphate; (*d*) iron(II) sulfate; (*e*) silver carbonate.
12. (*a*) $NaHSO_4$; (*b*) $PbCrO_4$; (*c*) CuCl; (*d*) $Hg_2(NO_3)_2$; (*e*) FeO.
13. (*a*) potassium sulfate; (*b*) barium dichromate; (*c*) chromium(III) hydroxide; (*d*) lead(II) bromide; (*e*) mercury(II) iodide.
14. (*a*) $Al(OH)_3$; (*b*) Cu_2O; (*c*) $(NH_4)_2S$; (*d*) $Pb(C_2H_3O_2)_2$; (*e*) $FeBr_3$.
15. (*a*) $Mg(HCO_3)_2$; (*b*) Ag_2S; (*c*) K_2SO_3; (*d*) $Cr_2(SO_4)_3$; (*e*) Na_3PO_4.

Group B

16. Atoms of essentially constant average mass combine in definite proportions to form the particles of more complex substances. Hence substances have a fixed percentage composition by mass.
17. The Atomic theory provides for the possibility that atoms may combine in more than one way and in more than one proportion.
18. (*a*) Sulfur trioxide; (*b*) silicon tetrachloride; (*c*) phosphorus tribromide; (*d*) diarsenic pentoxide; (*e*) lead monoxide.
19. (*a*) SO_2; (*b*) $BiCl_3$; (*c*) MnO_2; (*d*) AsI_5; (*e*) CI_4.
20. (*a*) Carbon monoxide; (*b*) carbon dioxide; (*c*) carbon tetrabromide; (*d*) dinitrogen trioxide; (*e*) dinitrogen pentoxide.

Answers to Problems pp. 140–142

Group A

1. Formula weight of $N_2H_4 = 2(14.0) + 4(1.0) = 32.0$
2. Formula weight of $H_2SO_4 = 2(1.0) + 32.1 + 4(16.0) = 98.1$
3. Formula weight of $C_6H_{12}O_6 = 6(12.0) + 12(1.0) + 6(16.0) = 180.0$
4. Formula weight of $C_2H_5OH = 2(12.0) + 6(1.0) + 16.0 = 46.0$
5. Formula weight of $Ca_3(PO_4)_2 = 3(40.1) + 2(31.0) + 8(16.0) = 310.3$
6. Formula weight of $MgSO_4 \cdot 7H_2O = 24.3 + 32.1 + 4(16.0) + 14(1.0) + 7(16.0) = 246.4$
7. (*a*) Formula weight of $HNO_3 = 1.0 + 14.0 + 3(16.0) = 63.0$
   (*b*) Formula weight of $NaOH = 23.0 + 16.0 + 1.0 = 40.0$
   (*c*) Formula weight of $HgO = 200.6 + 16.0 = 216.6$
   (*d*) Formula weight of $CuSO_4 \cdot 5H_2O = 63.5 + 32.1 + 4(16.0) + 10(1.0) + 5(16.0) = 249.6$
   (*e*) Formula weight of $HC_2H_3O_2 = 4(1.0) + 2(12.0) + 2(16.0) = 60.0$
   (*f*) Formula weight of $MgBr_2 = 24.3 + 2(79.9) = 184.1$
   (*g*) Formula weight of $Al_2S_3 = 2(27.0) + 3(32.1) = 150.3$
   (*h*) Formula weight of $Ca(NO_3)_2 = 40.1 + 2(14.0) + 6(16.0) = 164.1$
   (*i*) Formula weight of $Fe_2(Cr_2O_7)_3 = 2(55.8) + 6(52.0) + 21(16.0) = 759.6$
   (*j*) Formula weight of $KMnO_4 = 39.1 + 54.9 + 4(16.0) = 158.0$

8. Formula weight of $HC_2H_3O_2 = 4(1.0) + 2(12.0) + 2(16.0) = 60.0$

$$\text{H: } \frac{(4.0)\ \text{H}}{(60.0)\ HC_2H_3O_2} \times 100\%\ HC_2H_3O_2 = 6.7\%\ \text{H}$$

$$\text{C: } \frac{(24.0)\ \text{C}}{(60.0)\ HC_2H_3O_2} \times 100\%\ HC_2H_3O_2 = 40.0\%\ \text{C}$$

$$\text{O: } \frac{(32.0)\ \text{O}}{(60.0)\ HC_2H_3O_2} \times 100\%\ HC_2H_3O_2 = 53.3\%\ \text{O}$$

9. Formula weight of $NaHCO_3 = 23.0 + 1.0 + 12.0 + 3(16.0) = 84.0$

$$\text{Na: } \frac{(23.0)\ \text{Na}}{(84.0)\ NaHCO_3} \times 100\%\ NaHCO_3 = 27.4\%\ \text{Na}$$

$$\text{H: } \frac{(1.0)\ \text{H}}{(84.0)\ NaHCO_3} \times 100\%\ NaHCO_3 = 1.2\%\ \text{H}$$

$$\text{C: } \frac{(12.0)\ \text{C}}{(84.0)\ NaHCO_3} \times 100\%\ NaHCO_3 = 14.3\%\ \text{C}$$

$$\text{O: } \frac{(48.0)\ \text{O}}{(84.0)\ NaHCO_3} \times 100\%\ NaHCO_3 = 57.1\%\ \text{O}$$

10. Formula weight of $C_{17}H_{35}COONa = 18(12.0) + 35(1.0) + 2(16.0) + 23.0 = 306.0$

$$\text{C: } \frac{(216.0)\ \text{C}}{(306.0)\ C_{17}H_{35}COONa} \times 100\%\ C_{17}H_{35}COONa = 70.6\%\ \text{C}$$

$$\text{H: } \frac{(35.0)\ \text{H}}{(306.0)\ C_{17}H_{35}COONa} \times 100\%\ C_{17}H_{35}COONa = 11.4\%\ \text{H}$$

$$\text{O: } \frac{(32.0)\ \text{O}}{(306.0)\ C_{17}H_{35}COONa} \times 100\%\ C_{17}H_{35}COONa = 10.5\%\ \text{O}$$

$$\text{Na: } \frac{(23.0)\ \text{Na}}{(306.0)\ C_{17}H_{35}COONa} \times 100\%\ C_{17}H_{35}COONa = 7.5\%\ \text{Na}$$

11. (*a*) Formula weight of $SO_2 = 32.1 + 2(16.0) = 64.1$

$$\text{S: } \frac{(32.1)\ \text{S}}{(64.1)\ SO_2} \times 100\%\ SO_2 = 50.1\%\ \text{S}$$

$$\text{O: } \frac{(32.0)\ \text{O}}{(64.1)\ SO_2} \times 100\%\ SO_2 = 49.9\%\ \text{O}$$

(*b*) Formula weight of $Ca(OH)_2 = 40.1 + 2(16.0) + 2(1.0) = 74.1$

$$\text{Ca: } \frac{(40.1)\ \text{Ca}}{(74.1)\ Ca(OH)_2} \times 100\%\ Ca(OH)_2 = 54.1\%\ \text{Ca}$$

O: $\frac{(32.0)\ \text{O}}{(74.1)\ \text{Ca(OH)}_2} \times 100\%\ \text{Ca(OH)}_2 = 43.2\%\ \text{O}$

H: $\frac{(2.0)\ \text{H}}{(74.1)\ \text{Ca(OH)}_2} \times 100\%\ \text{Ca(OH)}_2 = 2.7\%\ \text{H}$

(*c*) Formula weight of $Ca(H_2PO_4)_2 \cdot H_2O = 40.1 + 4(1.0) + 2(31.0) + 8(16.0) + 18.0 = 252.1$

Ca: $\frac{(40.1)\ \text{Ca}}{(252.1)\ \text{Ca(H}_2\text{PO}_4)_2 \cdot \text{H}_2\text{O}} \times 100\%\ \text{Ca(H}_2\text{PO}_4)_2 \cdot \text{H}_2\text{O} = 15.9\%\ \text{Ca}$

H: $\frac{(4.0)\ \text{H}}{(252.1)\ \text{Ca(H}_2\text{PO}_4)_2 \cdot \text{H}_2\text{O}} \times 100\%\ \text{Ca(H}_2\text{PO}_4)_2 \cdot \text{H}_2\text{O} = 1.6\%\ \text{H}$

P: $\frac{(62.0)\ \text{P}}{(252.1)\ \text{Ca(H}_2\text{PO}_4)_2 \cdot \text{H}_2\text{O}} \times 100\%\ \text{Ca(H}_2\text{PO}_4)_2 \cdot \text{H}_2\text{O} = 24.6\%\ \text{P}$

O: $\frac{(128.0)\ \text{O}}{(252.1)\ \text{Ca(H}_2\text{PO}_4)_2 \cdot \text{H}_2\text{O}} \times 100\%\ \text{Ca(H}_2\text{PO}_4)_2 \cdot \text{H}_2\text{O} = 50.8\%\ \text{O}$

H_2O: $\frac{(18.0)\ \text{H}_2\text{O}}{(252.1)\ \text{Ca(H}_2\text{PO}_4)_2 \cdot \text{H}_2\text{O}} \times 100\%\ \text{Ca(H}_2\text{PO}_4)_2 \cdot \text{H}_2\text{O} = 7.1\%\ \text{H}_2\text{O}$

(*d*) Formula weight of $MgSO_4 \cdot 7H_2O = 24.3 + 32.1 + 4(16.0) + 7(18.0) = 246.4$

Mg: $\frac{(24.3)\ \text{Mg}}{(246.4)\ \text{MgSO}_4 \cdot 7\text{H}_2\text{O}} \times 100\%\ \text{MgSO}_4 \cdot 7\text{H}_2\text{O} = 9.9\%\ \text{Mg}$

S: $\frac{(32.1)\ \text{S}}{(246.4)\ \text{MgSO}_4 \cdot 7\text{H}_2\text{O}} \times 100\%\ \text{MgSO}_4 \cdot 7\text{H}_2\text{O} = 13.0\%\ \text{S}$

O: $\frac{(64.0)\ \text{O}}{(246.4)\ \text{MgSO}_4 \cdot 7\text{H}_2\text{O}} \times 100\%\ \text{MgSO}_4 \cdot 7\text{H}_2\text{O} = 26.0\%\ \text{O}$

H_2O: $\frac{(126.0)\ \text{H}_2\text{O}}{(246.4)\ \text{MgSO}_4 \cdot 7\text{H}_2\text{O}} \times 100\%\ \text{MgSO}_4 \cdot 7\text{H}_2\text{O} = 51.1\%\ \text{H}_2\text{O}$

12. (*a*) Formula weight of $Ca(NO_3)_2 = 40.1 + 2(14.0) + 6(16.0) = 164.1$

N: $\frac{(28.0)\ \text{N}}{(164.1)\ \text{Ca(NO}_3)_2} \times 100\%\ \text{Ca(NO}_3)_2 = 17.1\%\ \text{N}$

(*b*) Formula weight of $CaCN_2 = 40.1 + 12.0 + 2(14.0) = 80.1$

N: $\frac{(28.0)\ \text{N}}{(80.1)\ \text{CaCN}_2} \times 100\%\ \text{CaCN}_2 = 35.0\%\ \text{N}$

(*c*) Formula weight of $(NH_4)_2SO_4 = 2(14.0) + 8(1.0) + 32.1 + 4(16.0) = 132.1$

N: $\frac{(28.0)\ \text{N}}{(132.1)\ (\text{NH}_4)_2\text{SO}_4} \times 100\%\ (\text{NH}_4)_2\text{SO}_4 = 21.2\%\ \text{N}$

13. Mass of oxygen = 7.956 g – 6.356 g = 1.600 g

$$\text{Cu: } \frac{6.356 \text{ g Cu}}{7.956 \text{ g Cpd}} \times 100\% \text{ Cpd} = 78.89\% \text{ Cu}$$

$$\text{O: } \frac{1.600 \text{ g O}}{7.956 \text{ g Cpd}} \times 100\% \text{ Cpd} = 20.11\% \text{ O}$$

14. (*a*) 1.00 mole × 35.5 g Cl/mole = 35.5 g Cl
(*b*) 5.00 moles × 14.0 g N/mole = 70.0 g N
(*c*) 3.00 moles × 159.8 g Br_2/mole = 479 g Br_2
(*d*) 6.00 moles × 36.5 g HCl/mole = 219 g HCl
(*e*) 10.00 moles × 120.4 g $MgSO_4$/mole = 1204 g $MgSO_4$

15. (*a*) $\frac{25.00 \text{ g CaO}}{56.1 \text{ g/mole}} = 0.446$ mole CaO

(*b*) $\frac{25.00 \text{ g } Na_2CO_3 \cdot 10H_2O}{286.0 \text{ g/mole}} = 0.08741$ mole $Na_2CO_3 \cdot 10H_2O$

(*c*) $\frac{25.00 \text{ g } BaCl_2 \cdot 2H_2O}{244.3 \text{ g/mole}} = 0.1023$ mole $BaCl_2 \cdot 2H_2O$

(*d*) $\frac{25.00 \text{ g } (NH_4)_2SO_4}{132.1 \text{ g/mole}} = 0.1893$ mole $(NH_4)_2SO_4$

(*e*) $\frac{25.00 \text{ g } Fe(NO_3)_3 \cdot 6H_2O}{349.8 \text{ g/mole}} = 0.07147$ mole $Fe(NO_3)_3 \cdot 6H_2O$

16. Formula weight of Fe_3O_4 = 3(55.8) + 4(16.0) = 231.4
1.000 metric ton × 10^6 g/metric ton × 1 mole Fe_3O_4/231.4 g × 3 moles Fe/1 mole Fe_3O_4 = 1.296 × 10^4 moles Fe

17. Formula weight of HgS = 200.6 + 32.1 = 232.7
1.00 kg × 10^3 g/kg × 1 mole HgS/232.7 g × 1 mole Hg/1 mole HgS = 4.297 moles Hg

18. Formula weight of Cu_2O = 2(63.5) + 16.0 = 143.0

$$\text{Cu: } \frac{(127.0) \text{ Cu}}{(143.0) \text{ } Cu_2O} \times 100\% \text{ } Cu_2O = 88.81\% \text{ Cu}$$

Formula weight of $CuCO_3 \cdot Cu(OH)_2$ = 2(63.5) + 12.0 + 5(16.0) + 2(1.0) = 221.0

$$\text{Cu: } \frac{(127.0) \text{ Cu}}{(221.0) \text{ } CuCO_3 \cdot Cu(OH)_2} \times 100\% \text{ } CuCO_3 \cdot Cu(OH)_2 = 57.47\% \text{ Cu}$$

Formula weight of $CuFe_2S_4$ = 63.5 + 2(55.8) + 4(32.1) = 303.5

$$\text{Cu: } \frac{(63.5) \text{ Cu}}{(303.5) \text{ } CuFe_2S_4} \times 100\% \text{ } CuFe_2S_4 = 20.9\% \text{ Cu}$$

19. Formula weight of $CaCO_3$ = 40.1 + 12.0 + 3(16.0) = 100.1
Formula weight of CaO = 40.1 + 16.0 = 56.1

$$\frac{(56.1) \text{ CaO}}{(100.1) \text{ } CaCO_3} \times 100\% \text{ } CaCO_3 = 56.0\% \text{ CaO}$$

20. Formula weight of $CuSO_4 \cdot 5H_2O = 63.5 + 32.1 + 4(16.0) + 5(18.0) = 249.6$
Formula weight of $5H_2O = 5(18.0) = 90.0$

$$\frac{(90.0)\ H_2O}{(249.6)\ CuSO_4 \cdot 5H_2O} \times 100\%\ CuSO_4 \cdot 5H_2O = 36.1\%\ H_2O$$

Group B

21. % Pt = 100.0% – 42.1% = 57.9%

$$\text{Pt: } \frac{57.9\text{ g Pt}}{195.1\text{ g/mole}} = 0.297\text{ mole Pt}$$

$$\text{Cl: } \frac{42.1\text{ g Cl}}{35.5\text{ g/mole}} = 1.19\text{ moles Cl}$$

Pt:Cl = 0.297/0.297:1.19/0.297 = 1:4
Empirical formula = $PtCl_4$
% Pt = 100.0% – 26.7% = 73.3%

$$\text{Pt: } \frac{73.3\text{ g Pt}}{195.1\text{ g/mole}} = 0.376\text{ mole Pt}$$

$$\text{Cl: } \frac{26.7\text{ g Cl}}{35.5\text{ g/mole}} = 0.752\text{ mole Cl}$$

Pt:Cl = 0.376/0.376:0.752/0.376 = 1:2
Empirical formula = $PtCl_2$

22. % F = 100% – 85% = 15%

$$\text{Ag: } \frac{85\text{ g Ag}}{107.9\text{ g/mole}} = 0.79\text{ mole Ag}$$

$$\text{F: } \frac{15\text{ g F}}{19.0\text{ g/mole}} = 0.79\text{ mole F}$$

Ag:F = 1:1 Empirical formula = AgF

23. Formula weight of $C_{11}H_{12}N_2O_5Cl_2 = 11(12.0) + 12(1.0) + 2(14.0) + 5(16.0) + 2(35.5) = 323.0$

$$\text{C: } \frac{132.0\text{ C}}{323.0\ C_{11}H_{12}N_2O_5Cl_2} \times 100\%\ C_{11}H_{12}N_2O_5Cl_2 = 40.9\%\text{ C}$$

$$\text{H: } \frac{12.0\text{ H}}{323.0\ C_{11}H_{12}N_2O_5Cl_2} \times 100\%\ C_{11}H_{12}N_2O_5Cl_2 = 3.7\%\text{ H}$$

$$\text{N: } \frac{28.0\text{ N}}{323.0\ C_{11}H_{12}N_2O_5Cl_2} \times 100\%\ C_{11}H_{12}N_2O_5Cl_2 = 8.7\%\text{ N}$$

$$\text{O: } \frac{80.0\text{ O}}{323.0\ C_{11}H_{12}N_2O_5Cl_2} \times 100\%\ C_{11}H_{12}N_2O_5Cl_2 = 24.8\%\text{ O}$$

$$\text{Cl: } \frac{71.0\text{ Cl}}{323.0\ C_{11}H_{12}N_2O_5Cl_2} \times 100\%\ C_{11}H_{12}N_2O_5Cl_2 = 22.0\%\text{ Cl}$$

24. P: $\frac{43.67\text{ g P}}{31.0\text{ g/mole}} = 1.41\text{ moles P}$

O: $\frac{56.33\text{ g O}}{16.0\text{ g/mole}} = 3.52\text{ moles O}$

P:O = 1.41/1.41:3.52/1.41 = 1.00:2.50 = 2.00:5.00
Empirical formula = P_2O_5

25. K: $\frac{24.58\text{ gK}}{39.1\text{ g/mole}} = 0.629\text{ mole K}$

Mn: $\frac{34.81\text{ g Mn}}{54.9\text{ g/mole}} = 0.634\text{ mole Mn}$

O: $\frac{40.50\text{ gO}}{16.0\text{ g/mole}} = 2.53\text{ moles O}$

K:Mn:O = 0.629/0.629:0.634/0.629:2.53/0.629 = 1.00:1.01:4.02
Empirical formula = $KMnO_4$

26. Na: $\frac{37.70\text{ g Na}}{23.0\text{ g/mole}} = 1.64\text{ moles Na}$

Si: $\frac{22.95\text{ g Si}}{28.1\text{ g/mole}} = 0.817\text{ mole Si}$

O: $\frac{39.34\text{ g O}}{16.0\text{ g/mole}} = 2.46\text{ moles O}$

Na:Si:O = 1.64/0.817:0.817/0.817:2.46/0.817 = 2.01:1.00:3.01
Empirical formula = Na_2SiO_3

27. Na: $\frac{28.05\text{ g Na}}{23.0\text{ g/mole}} = 1.22\text{ moles Na}$

C: $\frac{29.26\text{ g C}}{12.0\text{ g/mole}} = 2.44\text{ moles C}$

H: $\frac{3.66\text{ g H}}{1.01\text{ g/mole}} = 3.62\text{ moles H}$

O: $\frac{39.02\text{ g O}}{16.0\text{ g/mole}} = 2.44\text{ moles O}$

Na:C:H:O = 1.22/1.22:2.44/1.22:3.62/1.22:2.44/1.22 = 1.00:2.00:2.97:2.00
Empirical formula = $NaC_2H_3O_2$

28. N: $\frac{21.20\text{ g N}}{14.0\text{ g/mole}} = 1.51\text{ moles N}$

H: $\frac{6.06\text{ g H}}{1.01\text{ g/mole}} = 6.00\text{ moles H}$

S: $\frac{24.30 \text{ g S}}{32.1 \text{ g/mole}} = 0.757 \text{ mole S}$

O: $\frac{48.45 \text{ g O}}{16.0 \text{ g/mole}} = 3.03 \text{ moles O}$

N:H:S:O = 1.51/0.757:6.00/0.757:0.757/0.757:3.03/0.757 = 1.99:7.93:1.00:4.00
Empirical formula = $N_2H_8SO_4$ or $(NH_4)_2SO_4$

29. K: $\frac{44.82 \text{ g K}}{39.1 \text{ g/mole}} = 1.15 \text{ moles K}$

S: $\frac{18.39 \text{ g S}}{32.1 \text{ g/mole}} = 0.573 \text{ mole S}$

O: $\frac{36.79 \text{ g O}}{16.0 \text{ g/mole}} = 2.30 \text{ moles O}$

K:S:O = 1.15/0.573:0.573/0.573:2.30/0.573 = 2.01:1.00:4.01
Empirical formula = K_2SO_4

30. Ca: $\frac{24.7 \text{ g Ca}}{40.1 \text{ g/mole}} = 0.616 \text{ mole Ca}$

H: $\frac{1.2 \text{ g H}}{1.0 \text{ g/mole}} = 1.2 \text{ moles H}$

C: $\frac{14.8 \text{ g C}}{12.0 \text{ g/mole}} = 1.23 \text{ moles C}$

O: $\frac{59.3 \text{ g O}}{16.0 \text{ g/mole}} = 3.71 \text{ moles O}$

Ca:H:C:O = 0.616/0.616:1.2/0.616:1.23/0.616:3.71/0.616 = 1.00:1.9:2.00:6.02
Empirical formula = $CaH_2C_2O_6$ or $Ca(HCO_3)_2$

31. Fe: $\frac{72.4 \text{ g Fe}}{55.8 \text{ g/mole}} = 1.30 \text{ moles Fe}$

O: $\frac{27.6 \text{ g O}}{16.0 \text{ g/mole}} = 1.72 \text{ moles O}$

Fe:O = 1.30/1.30:1.72/1.30 = 1.00:1.32 = 3.00:3.96
Empirical formula = Fe_3O_4

32. C: $\frac{92.3 \text{ g C}}{12.0 \text{ g/mole}} = 7.7 \text{ moles C}$

H: $\frac{7.7 \text{ g H}}{1.0 \text{ g/mole}} = 7.7 \text{ moles H}$

C:H = 1:1
Empirical formula = CH
$(\text{CH weight})_x = (12.0 + 1.0)_x = 26.0 \qquad x = 2$
Molecular formula = $(CH)_2$ or C_2H_2

33. C: $\frac{8\bar{0}\text{ g C}}{12\text{ g/mole}} = 6.7\text{ moles C}$

H: $\frac{2\bar{0}\text{ g H}}{1.0\text{ g/mole}} = 2\bar{0}\text{ moles H}$

C:H = 6.7/6.7:2$\bar{0}$/6.7 = 1.0:3.0
Empirical formula = CH_3
$(CH_3\text{ weight})_x = (12.0 + 3.0)_x = 30.0 \qquad x = 2$
Molecular formula = $(CH_3)_2$ or C_2H_6

34. % O = 100.0% – 42.8% = 57.2%

C: $\frac{42.8\text{ g C}}{12.0\text{ g/mole}} = 3.57\text{ moles C}$

O: $\frac{57.2\text{ g O}}{16.0\text{ g/mole}} = 3.58\text{ moles O}$

C:O = 1.1
% O = 100.0% – 27.3% = 72.7%

C: $\frac{27.3\text{ g C}}{12.0\text{ g/mole}} = 2.28\text{ moles C}$

O: $\frac{72.7\text{ g O}}{16.0\text{ g/mole}} = 4.54\text{ moles O}$

C:O = 2.28/2.28:4.54/2.28 = 1.00:1.99
In the two oxides, if C is constant, 1:1, then O is 1:1.99 or 1:2.

Chapter 8

Answers to Equations pp. 159–161

Group A

1. $Zn + S \rightarrow ZnS$. Composition.
2. $KCl + AgNO_3 \rightarrow AgCl(s) + KNO_3$. Ionic, precipitate.
3. $CaO + H_2O \rightarrow Ca(OH)_2$. Composition.
4. $NaOH + HCl \rightarrow NaCl + H_2O$. Ionic, molecular product (H_2O).
5. $MgBr_2 + Cl_2 \rightarrow MgCl_2 + Br_2$. Replacement of halogens.
6. $2NaCl + H_2SO_4 \rightarrow Na_2SO_4 + 2HCl(g)$. Ionic, gaseous product.
7. $2Al + Fe_2O_3 \rightarrow Al_2O_3 + 2Fe$. Replacement of a metal by a more active metal.
8. $NH_4NO_2 \rightarrow N_2(g) + 2H_2O$. Decomposition.
9. $2AgNO_3 + Cu \rightarrow Cu(NO_3)_2 + 2Ag(s)$. Replacement of a metal by a more active metal.
10. $3H_2 + N_2 \rightarrow 2NH_3(g)$. Composition.
11. sodium + iodine → sodium iodide. $2Na + I_2 \rightarrow 2NaI$.
12. magnesium + oxygen → magnesium oxide. $Mg + O_2 \rightarrow 2MgO$.
13. hydrogen + chlorine → hydrogen chloride(g). $H_2 + Cl_2 \rightarrow 2HCl(g)$. Reactions 11, 12, and 13 are exothermic (Chapter 6). The products are more stable than the reactants.
14. nickel(II) chlorate → nickel(II) chloride + oxygen(g). $Ni(ClO_3)_2 \rightarrow NiCl_2 + 3O_2(g)$. Decomposition of a metallic chlorate.
15. barium carbonate → barium oxide + carbon dioxide(g). $BaCO_3 \rightarrow BaO + CO_2(g)$. Decomposition of a metallic carbonate.
16. zinc hydroxide → zinc oxide + steam(g). $Zn(OH)_2 \rightarrow ZnO + H_2O(g)$. Decomposition of a metallic hydroxide.
17. aluminum + sulfuric acid → aluminum sulfate + hydrogen(g). $2Al + 3H_2SO_4 \rightarrow Al_2(SO_4)_3 + 3H_2(g)$. Aluminum is more active than hydrogen.

18. potassium iodide + chlorine → potassium chloride + iodine. $2KI + Cl_2 \rightarrow 2KCl + I_2$. Chlorine is more active than iodine.
19. iron + copper(II) nitrate → iron(II) nitrate + copper(s). $Fe + Cu(NO_3)_2 \rightarrow Fe(NO_3)_2 + Cu(s)$. Iron is more active than copper.
20. silver nitrate + zinc chloride → silver chloride(s) + zinc nitrate. $2AgNO_3 + ZnCl_2 \rightarrow 2AgCl(s) + Zn(NO_3)_2$. Silver chloride precipitates.
21. copper(II) hydroxide + acetic acid → copper(II) acetate + water. $Cu(OH)_2 + 2HC_2H_3O_2 \rightarrow Cu(C_2H_3O_2)_2 + 2H_2O$. Water is a molecular product.
22. iron(II) sulfate + ammonium sulfide → iron(II) sulfide(s) + ammonium sulfate. $FeSO_4 + (NH_4)_2S \rightarrow FeS(s) + (NH_4)_2SO_4$. Iron(II) sulfide precipitates.

Group B

23. barium chloride + sodium sulfate → barium sulfate(s) + sodium chloride. $BaCl_2 + Na_2SO_4 \rightarrow BaSO_4(s) + 2NaCl$. Ionic. Barium sulfate precipitates.
24. calcium + hydrochloric acid → calcium chloride + hydrogen(g). $Ca + 2HCl \rightarrow CaCl_2 + H_2(g)$. Replacement. Calcium is more active than hydrogen.
25. iron(II) sulfide + hydrochloric acid → hydrogen sulfide(g) + iron(II) chloride. $FeS + 2HCl \rightarrow H_2S(g) + FeCl_2$. Ionic. Hydrogen sulfide is a gas.
26. zinc chloride + ammonium sulfide → zinc sulfide(s) + ammonium chloride. $ZnCl_2 + (NH_4)_2S \rightarrow ZnS(s) + 2NH_4Cl$. Ionic. Zinc sulfide precipitates.
27. $NH_3 + 2O_2 \rightarrow HNO_3 + H_2O$.
28. magnesium + nitric acid → magnesium nitrate + hydrogen(g). $Mg + 2HNO_3 \rightarrow Mg(NO_3)_2 + H_2(g)$. Replacement. Magnesium is more active than hydrogen.
29. potassium + water → potassium hydroxide + hydrogen(g). $2K + 2H_2O \rightarrow 2KOH + H_2(g)$. Replacement. Potassium is more active than hydrogen.
30. sodium iodide + bromine → sodium bromide + iodine. $2NaI + Br_2 \rightarrow 2NaBr + I_2$. Replacement. Bromine is more active than iodine.
31. silver + sulfur → silver sulfide. $2Ag + S \rightarrow Ag_2S$. Composition. Exothermic reaction. Silver sulfide is more stable than elementary silver and sulfur.
32. sodium chlorate → sodium chloride + oxygen(g). $2NaClO_3 \rightarrow 2NaCl + 3O_2(g)$. Decomposition. Metallic chlorates yield metallic chlorides and oxygen.
33. $C + H_2O \rightarrow CO(g) + H_2(g)$.
34. zinc + lead(II) acetate → zinc acetate + lead(s). $Zn + Pb(C_2H_3O_2)_2 \rightarrow Zn(C_2H_3O_2)_2 + Pb(s)$. Replacement. Zinc is more active than lead.
35. iron(III) hydroxide → iron(III) oxide + steam. $2Fe(OH)_3 \rightarrow Fe_2O_3 + 3H_2O(g)$. Decomposition. Metallic hydroxides yield metallic oxides and steam.
36. $Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2(g)$.
37. lead(II) acetate + hydrogen sulfide → lead(II) sulfide(s) + acetic acid. $Pb(C_2H_3O_2)_2 + H_2S \rightarrow PbS(s) + 2HC_2H_3O_2$. Ionic. Lead(II) sulfide precipitates.
38. aluminum bromide + chlorine → aluminum chloride + bromine. $2AlBr_3 + 3Cl_2 \rightarrow 2AlCl_3 + 3Br_2$. Replacement. Chlorine is more active than bromine.
39. magnesium carbonate → magnesium oxide + carbon dioxide(g). $MgCO_3 \rightarrow MgO + CO_2(g)$. Decomposition. Metallic carbonates yield metallic oxides and carbon dioxide.
40. iron(III) chloride + sodium hydroxide → iron(III) hydroxide(s) + sodium chloride. $FeCl_3 + 3NaOH \rightarrow Fe(OH)_3(s) + 3NaCl$. Ionic. Iron(III) hydroxide precipitates.
41. $3CaO + P_2O_5 \rightarrow Ca_3(PO_4)_2$.
42. chromium + oxygen → chromium(III) oxide. $4Cr + 3O_2 \rightarrow 2Cr_2O_3$. Composition. Exothermic reaction. Chromium(III) oxide is more stable than elementary chromium and oxygen.
43. sodium + water → sodium hydroxide + hydrogen(s). $2Na + 2H_2O \rightarrow 2NaOH + H_2(g)$. Replacement. Sodium is more active than hydrogen.
44. calcium carbonate + hydrochloric acid → calcium chloride + carbon dioxide(g) + water. $CaCO_3 + 2HCl \rightarrow CaCl_2 + CO_2(g) + H_2O$. Ionic and decomposition of carbonic acid. Carbon dioxide is a gas.
45. calcium hydroxide + phosphoric acid → calcium phosphate + water. $3Ca(OH)_2 + 2H_3PO_4 \rightarrow Ca_3(PO_4)_2 + 6H_2O$. Ionic. Water is a molecular product.

46. sodium carbonate + nitric acid → sodium nitrate + carbon dioxide(g) + water. $Na_2CO_3 + 2HNO_3 \rightarrow 2NaNO_3 + CO_2(g) + H_2O$. Ionic and decomposition of carbonic acid. Carbon dioxide is a gas.
47. aluminum hydroxide + sulfuric acid → aluminum sulfate + water. $2Al(OH)_3 + 3H_2SO_4 \rightarrow Al_2(SO_4)_3 + 6H_2O$. Ionic. Water is a molecular product.
48. $Cu + 2H_2SO_4 \rightarrow CuSO_4 + 2H_2O + SO_2(g)$.
49. $Ca(OH)_2 + (NH_4)_2SO_4 \rightarrow CaSO_4 + 2NH_3(g) + 2H_2O$.
50 sodium sulfite + sulfuric acid → sodium sulfate + sulfur dioxide(g) + water. $Na_2SO_3 + H_2SO_4 \rightarrow Na_2SO_4 + SO_2(g) + H_2O$. Ionic and decomposition of sulfurous acid. Sulfur dioxide is a gas.

Solutions to Problems pp. 161-162

Group A

1. 25.0 g X
   $2HgO \rightarrow 2Hg + O_2(g)$
   2 moles 1 mole
   1 mole HgO = 216.6 g 1 mole O_2 = 32.0 g

   (*a*) $25.0 \text{ g HgO} \times \frac{\text{mole}}{216.6 \text{ g}} = 0.115 \text{ mole HgO}$

   (*b*) $0.115 \text{ mole HgO} \times \frac{1 \text{ mole } O_2}{2 \text{ moles HgO}} = 0.0575 \text{ mole } O_2$

   (*c*) $0.0575 \text{ mole } O_2 \times 32.0 \text{ g/mole} = 1.84 \text{ g } O_2$
2. 25.0 g X
   $2KClO_3 \rightarrow 2KCl + 3O_2(g)$
   2 moles 3 moles
   1 mole $KClO_3$ = 122.6 g 1 mole O_2 = 32.0 g

   (*a*) $25.0 \text{ g } KClO_3 \times \frac{\text{mole}}{122.6 \text{ g}} = 0.204 \text{ mole } KClO_3$

   (*b*) $0.204 \text{ mole } KClO_3 \times \frac{3 \text{ moles } O_2}{2 \text{ moles } KClO_3} = 0.306 \text{ mole } O_2$

   (*c*) $0.306 \text{ mole } O_2 \times 32.0 \text{ g/mole} = 9.79 \text{ g } O_2$
3. X 0.10 g
   $Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2(g)$
   1 mole 1 mole
   1 mole Zn = 65.4 g 1 mole H_2 = 2.0

   (*a*) $0.10 \text{ g } H_2 \times \frac{\text{mole}}{2.0 \text{ g}} = 0.050 \text{ mole } H_2$

   (*b*) $0.050 \text{ mole } H_2 \times \frac{1 \text{ mole Zn}}{1 \text{ mole } H_2} = 0.050 \text{ mole Zn}$

   (*c*) $0.050 \text{ mole Zn} \times 65.4 \text{ g/mole} = 3.3 \text{ g Zn}$

4. X 10.0 g
   $NaCl + AgNO_3 \rightarrow AgCl(s) + NaNO_3$
   1 mole 1 mole 1 mole
   1 mole $AgNO_3$ = 169.9 g 1 mole NaCl = 58.5 g

(*a*) $10.0 \text{ g } AgNO_3 \times \frac{\text{mole}}{169.9 \text{ g}} = 0.0589 \text{ mole } AgNO_3$

(*b*) $0.0589 \text{ mole } AgNO_3 \times \frac{1 \text{ mole NaCl}}{1 \text{ mole } AgNO_3} = 0.0589 \text{ mole NaCl}$

(*c*) $0.0589 \text{ mole NaCl} \times 58.5 \text{ g/mole} = 3.45 \text{ g NaCl}$

5. (*a*) $0.0589 \text{ mole } AgNO_3 \times \frac{1 \text{ mole AgCl}}{1 \text{ mole } AgNO_3} = 0.0589 \text{ mole AgCl}$

(*b*) 1 mole AgCl = 143.4 g
$0.0589 \text{ mole AgCl} \times 143.4 \text{ g/mole} = 8.45 \text{ g AgCl}$

6. $\underset{1 \text{ mole}}{S} + \underset{1 \text{ mole}}{O_2} \rightarrow \underset{}{SO_2(g)}$ (X over S, 80.0 g over SO_2)

1 mole SO_2 = 64.1 g 1 mole S = 32.1 g

$X = 80.0 \text{ g } SO_2 \times \frac{\text{mole}}{64.1 \text{ g}} \times \frac{1 \text{ mole S}}{1 \text{ mole } SO_2} \times 32.1 \text{ g/mole} = 40.1 \text{ g S}$

7. $\underset{1 \text{ mole}}{Fe_3O_4} + \underset{4 \text{ moles}}{4H_2} \rightarrow 3Fe + 4H_2O(g)$ (25 g over Fe_3O_4, X over $4H_2$)

1 mole Fe_3O_4 = 231.4 g 1 mole H_2 = 2.0 g

$X = 25 \text{ g } Fe_3O_4 \times \frac{\text{mole}}{231.4 \text{ g}} \times \frac{4 \text{ moles } H_2}{1 \text{ mole } Fe_3O_4} \times 2.0 \text{ g/mole} = 0.86 \text{ g } H_2$

8. $\underset{2 \text{ moles}}{2Cu} + \underset{2 \text{ moles}}{O_2} \rightarrow 2CuO$ (1.00 kg over Cu, X over O_2)

1 mole Cu = 63.5 g 1 mole CuO = 79.5 g

$X = 1.00 \text{ kg Cu} \times \frac{1000 \text{ g}}{\text{kg}} \times \frac{\text{mole}}{63.5 \text{ g}} \times \frac{2 \text{ moles CuO}}{2 \text{ moles Cu}} \times 79.5 \text{ g/mole}$

$= 1250 \text{ g CuO}$

9. $\underset{1 \text{ mole}}{Cu} + 2AgNO_3 \rightarrow \underset{2 \text{ moles}}{2Ag(s)} + Cu(NO_3)_2$ (40.0 g over Cu, X over 2Ag)

1 mole Cu = 63.5 g 1 mole Ag = 107.9 g

$X = 40.0 \text{ g Cu} \times \frac{\text{mole}}{63.5 \text{ g}} \times \frac{2 \text{ moles Ag}}{1 \text{ mole Cu}} \times 107.9 \text{ g/mole} = 136 \text{ g Ag}$

10. $\underset{1 \text{ mole}}{FeS} + 2HCl \rightarrow FeCl_2 + \underset{1 \text{ mole}}{H_2S(g)}$ (10.0 g over FeS, X over H_2S)

1 mole FeS = 87.9 g 1 mole H_2S = 34.1 g

$X = 10.0 \text{ g FeS} \times \frac{\text{mole}}{87.9 \text{ g}} \times \frac{1 \text{ mole } H_2S}{1 \text{ mole FeS}} \times 34.1 \text{ g/mole} = 3.88 \text{ g } H_2S$

Group B

11. $$\underset{1\text{ mole}}{H_2SO_4} + \underset{1\text{ mole}}{BaO_2} \rightarrow \underset{1\text{ mole}}{H_2O_2} + \underset{1\text{ mole}}{BaSO_4(s)}$$

$\overline{150}$ g X Y

1 mole BaO_2 = 169.3 g

(*a*) $$X = \overline{150}\text{ g } BaO_2 \times \frac{\text{mole}}{169.3\text{ g}} \times \frac{1\text{ mole } H_2O_2}{1\text{ mole } BaO_2} = 0.886\text{ mole } H_2O_2$$

(*b*) $$Y = 0.886\text{ mole } H_2O_2 \times \frac{1\text{ mole } BaSO_4}{1\text{ mole } H_2O_2} = 0.886\text{ mole } BaSO_4$$

12. $$\underset{2\text{ moles}}{Zn + 2HCl} \rightarrow \underset{1\text{ mole}}{ZnCl_2} + \underset{1\text{ mole}}{H_2(g)}$$

$\overline{100}$ g X

1 mole HCl = 36.5 g

$$X = \overline{100}\text{ g HCl} \times \frac{\text{mole}}{36.5\text{ g}} \times \frac{1\text{ mole } H_2}{2\text{ moles HCl}} = 1.37\text{ moles } H_2$$

13. $$\underset{1\text{ mole}}{Fe} + \underset{1\text{ mole}}{S} \rightarrow \underset{1\text{ mole}}{FeS}$$

10.0 g Y X

1 mole Fe = 55.8 g 1 mole S = 32.1 g 1 mole FeS = 87.9 g

(*a*) $$X = 10.0\text{ g Fe} \times \frac{\text{mole}}{55.8\text{ g}} \times \frac{1\text{ mole FeS}}{1\text{ mole Fe}} \times 8.79\text{ g/mole} = 15.8\text{ g FeS}$$

(*b*) $$Y = 10.0\text{ g Fe} \times \frac{\text{mole}}{55.8\text{ g}} \times \frac{1\text{ mole S}}{1\text{ mole Fe}} \times 32.1\text{ g/mole} = 5.75\text{ g S used}$$

10.0 g – 5.8 g = 4.2 g S remaining

$$\underset{1\text{ mole}}{S} + O_2 \rightarrow \underset{1\text{ mole}}{SO_2}$$

4.2 g X

1 mole S = 32.1 g 1 mole SO_2 = 64.1 g

$$X = 4.2\text{ g S} \times \frac{\text{mole}}{32.1\text{ g}} \times \frac{1\text{ mole } SO_2}{1\text{ mole S}} \times 64.1\text{ g/mole} = 8.4\text{ g } SO_2$$

14. $$CaCO_3 \rightarrow CaO + CO_2(g)$$

1 mole 1 mole

Equal mole quantities

$$\underset{1\text{ mole}}{CaO} + H_2O \rightarrow \underset{1\text{ mole}}{Ca(OH)_2}$$

X X

1 mole $CaCO_3$ = 100.1 g 1 mole $Ca(OH)_2$ = 74.1 g

$$X = 1.00 \text{ kg } CaCO_3 \times \frac{10^3 \text{ g}}{\text{kg}} \times \frac{\text{mole}}{100.1 \text{ g}} \times \frac{1 \text{ mole } Ca(OH)_2}{1 \text{ mole } CaCO_3} \times$$

$$74.1 \text{ g/mole} = 74\bar{0} \text{ g } Ca(OH)_2$$

15.
$$\begin{array}{ccccc} 93 \text{ g} & & X & & \\ 4P & + & 5O_2 & \rightarrow & P_4O_{10} \\ 4 \text{ moles} & & 5 \text{ moles} & & \end{array}$$

1 mole P = 31.0 g 1 mole O_2 = 32.0 g

$$X = 93 \text{ g P} \times \frac{\text{mole}}{31.0 \text{ g}} \times \frac{5 \text{ moles } O_2}{4 \text{ moles P}} \times 32.0 \text{ g/mole} \times \frac{100\% \text{ air}}{23\% \text{ } O_2} = 520 \text{ g air}$$

16.
$$\begin{array}{ccccc} & & X & & \\ C & + & O_2 & \rightarrow & CO_2(g) \\ 1 \text{ mole} & & & & 1 \text{ mole} \end{array}$$

1 mole C = 12.0 g 1 mole CO_2 = 44.0 g

$$X = 1.000 \text{ metric ton coke} \times \frac{9\bar{0}\% \text{ C}}{100\% \text{ coke}} \times \frac{\text{mole}}{12.0 \text{ g}} = \frac{1 \text{ mole } CO_2}{1 \text{ mole C}} \times$$

$$44.0 \text{ g/mole} = 3.3 \text{ metric tons } CO_2$$

17.
$$\begin{array}{ccccccc} & & X & & 0.50 \text{ mole} & & Y \\ 2Al & + & 3H_2SO_4 & \rightarrow & Al_2(SO_4)_3 & + & 3H_2(g) \\ & & 3 \text{ moles} & & 1 \text{ mole} & & 3 \text{ moles} \end{array}$$

1 mole H_2SO_4 = 98.1 g

(*a*) $$X = 0.50 \text{ mole } Al_2(SO_4)_3 \times \frac{3 \text{ moles } H_2SO_4}{1 \text{ mole } Al_2(SO_4)_3} \times 98.1 \text{ g/mole} = 150 \text{ g } H_2SO_4$$

(*b*) $$Y = 0.50 \text{ mole } Al_2(SO_4)_3 \times \frac{3 \text{ moles } H_2}{1 \text{ mole } Al_2(SO_4)_3} = 1.5 \text{ moles } H_2$$

18.
$$\begin{array}{ccccccc} 1.00 \text{ kg} & & X & & & & \\ 2C_4H_{10} & + & 13O_2 & \rightarrow & 8CO_2 & + & 10H_2O \\ 2 \text{ moles} & & 13 \text{ moles} & & & & \end{array}$$

1 mole C_4H_{10} = 58.0 g 1 mole O_2 = 32.0 g

$$X = 1.00 \text{ kg } C_4H_{10} \times \frac{\text{mole}}{58.0 \text{ g}} \times \frac{13 \text{ moles } O_2}{2 \text{ moles } C_4H_{10}} \times 32.0 \text{ g/mole} = 3.59 \text{ kg } O_2$$

19.
$$\begin{array}{ccccccc} 45 \text{ g} & & & & X & & Y \\ 2C_2H_6 & + & 7O_2 & \rightarrow & 4CO_2 & + & 6H_2O \\ 2 \text{ moles} & & & & 4 \text{ moles} & & 6 \text{ moles} \end{array}$$

1 mole C_2H_6 = 30.0 g

(*a*) $$X = 45 \text{ g } C_2H_6 \times \frac{\text{mole}}{30.0 \text{ g}} \times \frac{4 \text{ moles } CO_2}{2 \text{ moles } C_2H_6} = 3.0 \text{ moles } CO_2$$

(*b*) $$Y = 45 \text{ g } C_2H_6 \times \frac{\text{mole}}{30.0 \text{ g}} \times \frac{6 \text{ moles } H_2O}{2 \text{ moles } C_2H_6} = 4.5 \text{ moles } H_2O$$

20. $\underset{1\text{ mole}}{H_2SO_4} + \underset{2\text{ moles}}{2NaCl} \rightarrow \underset{1\text{ mole}}{Na_2SO_4} + \underset{2\text{ moles}}{2HCl}$ (150 g H_2SO_4; Y; X; Z)

1 mole H_2SO_4 = 98.1 g 1 mole NaCl = 58.5 g 1 mole Na_2SO_4 = 142.1 g
1 mole HCl = 36.5 g

(*a*) $X = 150\text{ g } H_2SO_4 \times \frac{\text{mole}}{98.1\text{ g}} \times \frac{1\text{ mole } Na_2SO_4}{1\text{ mole } H_2SO_4} \times 142.1\text{ g/mole} = 217\text{ g } Na_2SO_4$

(*b*) $Y = 150\text{ g } H_2SO_4 \times \frac{\text{mole}}{98.1\text{ g}} \times \frac{2\text{ moles NaCl}}{1\text{ mole } H_2SO_4} \times 58.5\text{ g/mole} = 179\text{ g NaCl}$

(*c*) $Z = 150\text{ g } H_2SO_4 \times \frac{\text{mole}}{98.1\text{ g}} \times \frac{2\text{ moles HCl}}{1\text{ mole } H_2SO_4} \times 36.5\text{ g/mole} = 112\text{ g HCl}$

Chapter 9

Answers to Questions pp. 178–179

Group A

1. Matter is composed of very tiny particles. The particles of matter are in continual motion. The particles of matter do not lose energy in collisions.
2. Because the collisions of the gas particles with each other and with the walls of the container are perfectly elastic.
3. (*a*) The higher the temperature of a gas, the more kinetic energy its particles have: the lower the temperature of a gas, the less kinetic energy its particles have. (*b*) No. The kinetic energy possessed by individual gas particles, all of which are at the same temperature, varies widely. However, most of the particles will have kinetic energies near the average kinetic energy of all the particles.
4. (*a*) The lowest temperature at which a substance may exist as a gas at atmospheric pressure. (*b*) The gas may condense to a liquid. (*c*) The kinetic energy of the gas particles is not sufficient to overcome the forces of attraction between them.
5. Van der Waals forces and the attraction between the oppositely charged portions of neighboring polar molecules. Van der Waals forces are close range attractive forces between molecules which depend on the number of electrons in a molecule and the tightness with which they are held. The greater the number of electrons and the less tightly they are bound, the stronger the Van der Waals forces are. The attraction between polar molecules is the electrostatic attraction of oppositely charged bodies.
6. It does not indicate the conditions under which the measurement is made.
7. (*a*) Exactly 0° C. (*b*) Exactly 760 mm of mercury
8. (*a*) The volume of a certain amount of dry gas is inversely proportional to the pressure, provided the temperature remains constant. (*b*) $V' = Vp/p'$.
9. (*a*) −273° C. (*b*) Kelvin temperature = Celsius temperature + 273°.
10. The volume of a certain amount of dry gas varies directly with the Kelvin temperature, provided the pressure remains constant. $V' = VT'/T$.
11. Its pressure may be greater than, less than, or the same as that of the air in the room.

Group B

12. (*a*) The particles of a gas move apart. (*b*) The particles of a gas bombard the walls of the container. (*c*) The particles of a gas are far apart. (*d*) The particles of a gas can move in and occupy truly empty space, or the spaces between other gas particles.
13. If each linear dimension is increased tenfold, the volume is increased 10^3 or 1000 times.
14. The condensation temperatures of polar covalent molecular substances are generally higher than those of nonpolar covalent molecular substances; hence, the strength of the attractive forces between particles of polar covalent molecular substances is greater than that between particles of nonpolar covalent molecular substances.

15. The condensation temperatures of ionic, covalent network, and metallic substances are generally higher than those of molecular substances; hence, the strength of the attractive forces between particles of ionic, covalent network, and metallic substances is greater than that between particles of molecular substances.
16. Pressure is directly proportional to the Kelvin temperature.
17. Volume is directly proportional to the Kelvin temperature.
18. Volume is inversely proportional to the pressure.
19. (*a*) The partial pressure exerted by water vapor. (*b*) It makes the pressure of the gas appear greater than it actually is. (*c*) By subtracting the water vapor pressure at the given temperature from the measured total pressure of the gas within the tube.
20. (*a*) No correction. (*b*) Subtract difference in levels from barometric pressure. (*c*) Subtract water vapor pressure. (*d*) Subtract water vapor pressure and difference in levels divided by 13.6.
21. At temperatures near room temperature and at pressures of less than a few atmospheres.

Solutions to Problems pp. 178--179

Group A

1. $V' = Vp/p'$. $V' = 25\bar{0}$ ml $\times$ $72\bar{0}$ mm/$75\bar{0}$ mm. $V' = 24\bar{0}$ ml.
2. $V' = Vp/p'$. $V' = 38\bar{0}$ ml $\times$ $80\bar{0}$ mm/$76\bar{0}$ mm. $V' = 40\bar{0}$ ml.
3. $V' = Vp/p'$. $V' = 10\bar{0}$ ml $\times$ 735 mm/$70\bar{0}$ mm. $V' = 105$ ml.
4. $V' = Vp/p'$. $p' = pV/V'$. $p' = 70.0$ cm $\times$ 240.0 ml/60.0 ml. $p' = 28\bar{0}$ cm.
5. (*a*) 293° K. (*b*) 358° K. (*c*) 258° K. (*d*) 83° K.
6. 27° C = $30\bar{0}$° K. 42° C = 315° K. $V' = VT'/T$. $T' = 90.0$ ml $\times$ 315° K/$30\bar{0}$° K. $V' =$ 94.5 ml.
7. 43° C = 316° K. $V' = VT'/T$. $T' = TV'/V$. $T' = 316$° K $\times$ 135 ml/$18\bar{0}$ ml. $T' =$ 237° K. 237° K = −36° C.
8. −23° C = $25\bar{0}$° K. 23° C = 296° K. $V' = VT'/T$. $V' = 50\bar{0}$ ml $\times$ 296° K/$25\bar{0}$° K. $V' =$ 592 ml.
9. 27° C = $30\bar{0}$° K. 0° C = 273° K. $V' = VT'/T$. $V' = 50.0$ liters $\times$ 273° K/$30\bar{0}$° K. $V' =$ 45.5 liters.
10. $3\bar{0}$° C = 303° K. $\bar{0}$° C = 273° K. $V' = V \times p/p' \times T'/T$. $V' = 2280$ ml $\times$ 808 mm/$76\bar{0}$ mm $\times$ 273° K/303° K. $V' = 2.18 \times 10^3$ ml.
11. $-2\bar{3}$° C = $25\bar{0}$° K. $\bar{0}$° C = 273° K. $V' = V \times p/p' \times T'/T$. $V' = 10\bar{0}0$ ml $\times$ $70\bar{0}$ mm/$76\bar{0}$ mm $\times$ $2\bar{7}3$° K/$25\bar{0}$° K. $V' = 1.01 \times 10^3$ ml.
12. $-3\bar{3}$° C = $24\bar{0}$° K. $\bar{0}$° C = 273° K. $V' = V \times p/p' \times T'/T$. $V' = 1520$ ml $\times$ $72\bar{0}$ mm/$76\bar{0}$ mm $\times$ $2\bar{7}3$° K/$24\bar{0}$° K. $V' = 1.64 \times 10^3$ ml.
13. 27° C = $30\bar{0}$° K. −3° C = $27\bar{0}$° K. $V' = V \times p/p' \times T'/T$. $V' = 50\bar{0}$ ml $\times$ 80.0 cm/75.0 cm $\times$ $2\bar{7}0$° K/300° K. $V' = 48\bar{0}$ ml.
14. 17° C = $29\bar{0}$° K. 307° C = $58\bar{0}$° K. $V' = V \times p/p' \times T'/T$. $V' = 10\bar{0}$ ml $\times$ $38\bar{0}$ mm/$50\bar{0}$ mm $\times$ $58\bar{0}$° K/$29\bar{0}$° K. $V' = 152$ ml.

Group B

15. 740.0 mm − $4\bar{0}$ mm = $70\bar{0}$ mm. $2\bar{5}$° C = 298° K. $\bar{0}$° C = 273° K. $V' = V \times p/p' \times T'/T$. $V' = 35.0$ ml $\times$ $70\bar{0}$ mm/$76\bar{0}$ mm $\times$ 273° K/298° K. $V' = 29.5$ ml
16. 715 mm − 25 mm = $69\bar{0}$ mm. 20° C = 293° K. $\bar{0}$° C = 273° K. $V' = V \times p/p' \times T'/T$. $V' = 60.0$ ml $\times$ $69\bar{0}$ mm/$76\bar{0}$ mm $\times$ 273° K/$29\bar{3}$° K. $V' = 50.8$ ml.
17. 720.0 mm − 14.5 mm = 705.5 mm. 17° C = $29\bar{0}$° K. $\bar{0}$° C = 273° K. $V' = V \times p/p' \times T'/T$. $V' = 25.0$ ml $\times$ 705.5 mm/760.0 mm $\times$ 273° K/$29\bar{0}$° K. $V' = 21.8$ ml.
18. 732.0 mm − 21.1 mm = 710.9 mm. 23° C = 296° K. $\bar{0}$° C = 273° K. $V' = V \times p/p' \times T'/T$. $V' = 45.0$ ml $\times$ 710.9 mm/760.0 mm $\times$ 273° K/296° K. $V' = 38.8$ ml.
19. 727.0 mm − (65 mm ÷ 13.6) = 722.2 mm. 722.2 mm − 23.8 mm = 698.4 mm. 25° C = 298° K. $\bar{0}$° C = 273° K. $V' = V \times p/p' \times T'/T$. $V' = 50.0$ ml $\times$ 698.4 mm/760.0 mm $\times$ 273° K/298° K. $V' = 42.1$ ml.
20. 745.0 mm − (95 mm ÷ 13.6) = 738.0 mm. 738.0 mm − 15.5 mm = 722.5 mm. 18° C = 291° K. $\bar{0}$° C = 273° K. $V' = V \times p/p' \times T'/T$. $V' = 12.0$ ml $\times$ 722.5 mm/760.0 mm $\times$ 273° K/291° K. $V' = 10.7$ ml.
21. 1.98 g/liter $\times$ ($76\bar{0}$ mm + $4\bar{0}$ mm)/$76\bar{0}$ mm = 2.08 g/liter.
22. $\bar{0}$° C = 273° K. 39° C = 312° K. 1.43 g/liter $\times$ 273° K/312° K = 1.25 g/liter.

23. $\bar{0}°$ C = 273° K. 27° C = $30\bar{0}°$ K. 1.25 g/liter × 273° K/$30\bar{0}°$ K × 90.0 cm/76.0 cm = 1.35 g/liter.
24. 25° C = 298° K. $V' = V \times p/p' \times T'/T$. $T' = T \times V'/V \times p'/p$. T' = 298° K × $35\bar{0}$ ml/$40\bar{0}$ ml × $74\bar{0}$ mm/$80\bar{0}$ mm. T' = 241° K. 241° K = −32° C.

Chapter 10
Solutions to Problems pp. 197-199

Group A

1. g-mol wt of HCl = (1.0 + 35.5)g = 36.5 g. D = g-mol wt/molar volume = 36.5 g/22.4 liters = 1.63 g/liter at STP.
2. g-mol wt of H_2S = 2(1.0)g + 32.1 g = 34.1 g. D = g-mol wt/molar volume = 34.1 g/22.4 liters = 1.52 g/liter at STP.
3. g-mol wt of CH_4 = 12.0 g + 4(1.0)g = 16.0 g. D = g-mol wt/molar volume = 16.0 g/22.4 liters = 0.714 g/liter at STP. The mass of 1.00 liter is 0.714 g.
4. g-mol wt = D × molar volume = 2.75 g/1.00 liter × 22.4 liters = 61.6 g. mol wt = 61.6.
5. (*a*) g-mol wt = D × molar volume = 1.25 g/1.00 liter × 22.4 liters = 28.0 g. mol wt = 28.0. (*b*) If the atomic weight of nitrogen is 14.0 and its molecular weight is 28.0, there must be 28.0/14.0 or 2.00 atoms/molecule.
6. g-mol wt of H_2 = 2(1.01)g = 2.02 g. $D = m/V$ = g-mol wt/molar volume. m = g-mol wt/molar volume × V = 2.02 g/22.4 liters × 0.300 liter = 0.0271 g, the mass of 300 ml of H_2 at STP.
7. m/V = g-mol wt/molar volume. g-mol wt = m/V × molar volume = 0.6428 g/0.225 liter × 22.4 liters = 64.0 g. mol wt = 64.0.
8. g-mol wt of CO_2 = 12.0 g + 2(16.0)g = 44.0 g. m/V = g-mol wt/molar volume. m = g-mol wt/molar volume × V = 44.0 g/22.4 liters × 0.750 liter = 1.47 g.
9. m/V = g-mol wt/molar volume. g-mol wt = m/V × molar volume = 0.179 g/0.250 liter × 22.4 liters = 16.0 g. mol wt = 16.0.
10. (*a*) mol wt of HBr = 1.0 + 79.9 = 80.9
    mol wt of PH_3 = 31.0 + 3(1.0) = 34.0
    mol wt of N_2O = 2(14.0) + 16.0 = 44.0
    (*b*) D of HBr = g-mol wt/molar volume = 80.9 g/22.4 liters = 3.61 g/liter
    D of PH_3 = g-mol wt/molar volume = 34.0 g/22.4 liters = 1.52 g/liter
    D of N_2O = g-mol wt/molar volume = 44.0 g/22.4 liters = 1.96 g/liter
11. m/V = g-mol wt/molar volume. m = g-mol wt/molar volume × V.
    N_2: m = 28.0 g/22.4 liters × 4.00 liters = 5.00 g
    NH_3: m = 17.0 g/22.4 liters × 4.00 liters = 3.04 g
    C_2H_2: m = 26.0 g/22.4 liters × 4.00 liters = 4.64 g

12. $$\begin{array}{ccccc} & Y & & X & 2\bar{0}\text{ liters} \\ N_2(g) & + & 3H_2(g) & \rightarrow & 2NH_3(g) \\ 1\text{ mole} & & 3\text{ moles} & & 2\text{ moles} \end{array}$$

$$X = 2\bar{0}\text{ liters } NH_3 \times \frac{3\text{ moles } H_2}{2\text{ moles } NH_3} = 3\bar{0}\text{ liters } H_2$$

$$Y = 2\bar{0}\text{ liters } NH_3 \times \frac{1\text{ mole } N_2}{2\text{ moles } NH_3} = 1\bar{0}\text{ liters } N_2$$

13. $$\begin{array}{ccccc} 15\text{ liters} & & Y & & X \\ 2CO(g) & + & O_2(g) & \rightarrow & 2CO_2(g) \\ 2\text{ moles} & & 1\text{ mole} & & 2\text{ moles} \end{array}$$

(*a*) $$X = 15\text{ liters CO} \times \frac{2\text{ moles } CO_2}{2\text{ moles CO}} = 15\text{ liters } CO_2$$

(*b*) $$Y = 15\text{ liters CO} \times \frac{1\text{ mole } O_2}{2\text{ moles CO}} = 7.5\text{ liters } O_2$$

14. 25.0 liters X Y
$2C_2H_2(g) + 5O_2(g) \rightarrow 4CO_2(g) + 2H_2O(l)$
2 moles 5 moles 4 moles

(*a*) $X = 25.0 \text{ liters } C_2H_2 \times \frac{5 \text{ moles } O_2}{2 \text{ moles } C_2H_2} = 62.5 \text{ liters } O_2$

(*b*) $Y = 25.0 \text{ liters } C_2H_2 \times \frac{4 \text{ moles } CO_2}{2 \text{ moles } C_2H_2} = 50.0 \text{ liters } CO_2$

15. 12 liters X(liters) Y(moles
$2C_2H_6(g) + 7O_2(g) \rightarrow 4CO_2(g) + 6H_2O(l)$
2 moles 7 moles 4 moles 6 moles

(*a*) $X = 12 \text{ liters } C_2H_6 \times \frac{4 \text{ moles } CO_2}{2 \text{ moles } C_2H_6} = 24 \text{ liters } CO_2$

(*b*) $Y = \frac{12 \text{ liters } C_2H_6}{22.4 \text{ liters/mole}} \times \frac{6 \text{ moles } H_2O}{2 \text{ moles } C_2H_6} = 1.6 \text{ moles } H_2$

16. $12 \text{ liters } C_2H_6 \times \frac{7 \text{ moles } O_2}{2 \text{ moles } C_2H_6} \times \frac{100\% \text{ air}}{21\% \ O_2} = 200 \text{ liters air}$

17. 400 ml X Y
$2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$
2 moles 1 mole 2 moles

(*a*) $X = 400 \text{ ml } H_2 \times \frac{1 \text{ mole } O_2}{2 \text{ moles } H_2} = 200 \text{ ml } O_2$ required

$400 \text{ ml} - 200 \text{ ml} = 200 \text{ ml } O_2$ uncombined

(*b*) $Y = 400 \text{ ml } H_2 \times \frac{2 \text{ moles } H_2O}{2 \text{ moles } H_2} = 400 \text{ ml } H_2O$ vapor

18. X(g) 4.0 liters
$2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$
2 moles 1 mole

1 mole Na = 23.0 g 1 mole H_2 = 22.4 liters

$X = 4.0 \text{ liters } H_2 \times \frac{\text{mole}}{22.4 \text{ liters}} \times \frac{2 \text{ moles Na}}{1 \text{ mole } H_2} \times 23.0 \text{ g/mole} = 8.2 \text{ g Na}$

19. 25.0 g X(liters) Y(moles)
$CuO(s) + H_2(g) \rightarrow Cu(s) + H_2O(l)$
1 mole 1 mole 1 mole

1 mole CuO = 79.5 g 1 mole H_2 = 22.4 liters

(*a*) $X = 25.0 \text{ g CuO} \times \frac{\text{mole}}{79.5 \text{ g}} \times \frac{1 \text{ mole } H_2}{1 \text{ mole CuO}} \times 22.4 \text{ liters/mole} = 7.04 \text{ liters } H_2$

(*b*) $Y = 25.0 \text{ g CuO} \times \frac{\text{mole}}{79.5 \text{ g}} \times \frac{1 \text{ mole } H_2O}{1 \text{ mole CuO}} = 0.314 \text{ mole } H_2O$

20. $\underset{1\text{ mole}}{Zn(s)} + \underset{2\text{ moles}}{2HCl(aq)} \rightarrow ZnCl_2(aq) + \underset{1\text{ mole}}{H_2(g)}$

1 mole Zn = 65.4 g 1 mole HCl = 36.5 g 1 mole H_2 = 22.4 liters

By inspection HCl is seen to be in excess.

$$X = 13\bar{0}\text{ g Zn} \times \frac{\text{mole}}{65.4\text{ g}} \times \frac{1\text{ mole } H_2}{1\text{ mole Zn}} \times 22.4\text{ liters/mole} = 44.5\text{ liters } H_2$$

21. $\underset{2\text{ moles}}{2H_2O(l)} \rightarrow \underset{1\text{ mole}}{O_2(g)} + \underset{2\text{ moles}}{2H_2(g)}$

1 mole H_2O = 18.0 g 1 mole O_2 = 22.4 liters 1 mole H_2 = 22.4 liters

(*a*) $$X = 90.0\text{ g } H_2O \times \frac{\text{mole}}{18.0\text{ g}} \times \frac{1\text{ mole } O_2}{2\text{ moles } H_2O} \times 22.4\text{ liters/mole} = 56.0\text{ liters } O_2$$

(*b*) $$Y = 90.0\text{ g } H_2O \times \frac{\text{mole}}{18.0\text{ g}} \times \frac{2\text{ moles } H_2}{2\text{ moles } H_2O} \times 22.4\text{ liters/mole} = 112\text{ liters } H_2$$

22. $\underset{1\text{ mole}}{CuO(s)} + \underset{1\text{ mole}}{H_2(g)} \rightarrow \underset{1\text{ mole}}{Cu(s)} + H_2O(l)$

1 mole CuO = 79.5 g 1 mole Cu = 63.5 g 1 mole H_2 = 22.4 liters

(*a*) $$X = 39.75\text{ g CuO} \times \frac{\text{mole}}{79.5\text{ g}} \times \frac{1\text{ mole Cu}}{1\text{ mole CuO}} \times 63.5\text{ g/mole} = 31.8\text{ g Cu}$$

(*b*) $$Y = 39.75\text{ g CuO} \times \frac{\text{mole}}{79.5\text{ g}} \times \frac{1\text{ mole } H_2}{1\text{ mole CuO}} \times 22.4\text{ liters/mole} = 11.2\text{ liters } H_2$$

23. $\underset{1\text{ mole}}{Ca(s)} + 2HCl(aq) \rightarrow CaCl_2(aq) + \underset{1\text{ mole}}{H_2(g)}$

1 mole Ca = 40.1 g 1 mole H_2 = 22.4 liters

$$X = 25\text{ g Ca} \times \frac{\text{mole}}{40.1\text{ g}} \times \frac{1\text{ mole } H_2}{1\text{ mole Ca}} \times 22.4\text{ liters/mole} = 14\text{ liters } H_2$$

Group B

24. (*a*) $$\frac{30.51\text{ g N}}{14.0\text{ g/mole}} = 2.18\text{ moles N}$$

$$\frac{69.49\text{ g O}}{16.0\text{ g/mole}} = 4.34\text{ moles O}$$

N:O = 2.18/2.18:4.34/2.18 = 1.00:1.99
Empirical formula = NO_2
(*b*) g-mol wt = $D \times$ molar volume = 4.085 g/liter × 22.4 liters = 91.5 g.
mol wt = 91.5.
(*c*) $(NO_2\text{ weight})_x = (14.0 + 32.0)_x = 91.5_x = 2.0$
Molecular formula = $(NO_2)_2$ or N_2O_4

25. g-mol wt = $m/V \times$ molar volume = 1.30 g/[1.00 liter $\times$ 72$\bar{0}$ mm/76$\bar{0}$ mm $\times$ 273° K/30$\bar{0}$° K] $\times$ 22.4 liters = 33.8 g. mol wt = 33.8.

26. (*a*) $\frac{46.7 \text{ g N}}{14.0 \text{ g/mole}} = 3.33$ moles N

$\frac{53.3 \text{ g O}}{16.0 \text{ g/mole}} = 3.33$ moles O

N:O = 3.33/3.33:3.33/3.33 = 1:1
Empirical formula = NO
(*b*) g-mol wt = $D \times$ molar volume = 1.34 g/liter $\times$ 22.4 liters = 30.0 g.
mol wt = 30.0. molecular formula = NO.
(*c*) Given: 1.00 liter of nitrogen + 1.00 liter of oxygen → 2.00 liters of NO. From Avogadro's principle: 1 molecule of nitrogen + 1 molecule of oxygen → 2 molecules of NO. Each of two molecules of NO contains one atom of N and one atom of O. Thus each nitrogen molecule must consist of two atoms and each oxygen molecule must consist of two atoms.

27. m/V = g-mol wt/molar volume. $V = m/$[g-mol wt/molar volume] = 2.0 g/[76.2 g/22.4 liters] = 0.59 liter at STP. $V_{50°\text{C, }756\text{ mm}}$ = 0.59 liter $\times$ 76$\bar{0}$ mm/756 mm $\times$ 323° K/273°K = 0.70 liter.

28. 6.02×10^{23} molecules $\times \frac{1.00 \text{ liter}}{22.4 \text{ liters}} \times \frac{1.00 \times 10^{-4} \text{ mm}}{76\bar{0} \text{ mm}} = 3.54 \times 10^{15}$ molecules.

29. g-mol wt = $m/V \times$ molar volume = 0.865 g/[0.174 liter $\times$ 273° K/373° K $\times$ 745 mm/76$\bar{0}$mm] $\times$ 22.4 liters = 155 g. mol wt = 155.

30. (*a*)
$$\underset{1\text{ mole}}{S(s)} + O_2(g) \rightarrow \underset{1\text{ mole}}{SO_2(g)}$$
(5$\bar{0}$ g above S; X above SO_2)

1 mole S = 32.1 g 1 mole SO_2 = 22.4 liters

$X = 5\bar{0}\text{ g S} \times \frac{\text{mole}}{32.1\text{ g}} \times \frac{1\text{ mole } SO_2}{1\text{ mole S}} \times 22.4\text{ liters/mole} = 35\text{ liters } SO_2 \text{ at STP.}$

(*b*) $V_{25°\text{ C, }745\text{ mm}} = 35\text{ liters} \times 76\bar{0}\text{ mm}/745\text{ mm} \times 298°\text{K}/273°\text{K} = 39\text{ liters.}$

31. $V_{STP} = 40\bar{0}\text{ ml} \times 74\bar{0}\text{ mm}/76\bar{0}\text{ mm}/ \times 273°\text{K}/293°\text{K} = 363\text{ ml.}$

$$\underset{1\text{ mole}}{Mg(s)} + 2HCl(aq) \rightarrow MgCl_2(aq) + \underset{1\text{ mole}}{H_2(g)}$$
(X above Mg; 0.363 liter above H_2)

1 mole H_2 = 22.4 liters 1 mole Mg = 24.3 g

$X = 0.363\text{ liter } H_2 \times \frac{\text{mole}}{22.4\text{ liter}} \times \frac{1\text{ mole Mg}}{1\text{ mole } H_2} \times 24.3\text{ g/mole} = 0.394\text{ g Mg}$

32.
$$\underset{1\text{ mole}}{Mg(s)} + H_2SO_4(aq) \rightarrow MgSO_4(aq) + \underset{1\text{ mole}}{H_2(g)}$$
(6.0 g above Mg; X above H_2)

1 mole Mg = 24.3 g 1 Mole H_2 = 22.4 liters

$$X = 6.0 \text{ g Mg} \times \frac{\text{mole}}{24.3 \text{ g}} \times \frac{1 \text{ mole } H_2}{1 \text{ mole Mg}} \times 22.4 \text{ liters/mole} = 5.5 \text{ liters } H_2 \text{ at STP.}$$

$V_{25^\circ \text{C}, 755.0 \text{ mm}} = 5.5 \text{ liters} \times 760 \text{ mm}/(755.0 \text{ mm} - 23.8 \text{ mm}) \times 298^\circ\text{K}/273^\circ\text{K} = 6.2 \text{ liters}$

33. $V_{STP} = 12.0 \text{ liters} \times (745.0 \text{ mm} - 21.1 \text{ mm})/760 \text{ mm} \times 273^\circ\text{K}/296^\circ\text{K} = 10.5 \text{ liters}$ at STP. $10.5 \text{ liters} \times 32.0 \text{ g}/22.4 \text{ liters} = 15.0 \text{ g}.$

34.

$$\underset{2 \text{ moles}}{2Al(s)} + 3H_2SO_4(aq) \rightarrow Al_2(SO_4)_3(aq) + \underset{3 \text{ moles}}{3H_2(g)}$$

1 mole Al = 27.0 g 1 mole H_2 = 22.4 liters

$$X = 50.0 \text{ g Al} \times \frac{\text{mole}}{27.0 \text{ g}} \times \frac{3 \text{ moles } H_2}{2 \text{ moles Al}} \times 22.4 \text{ liters/mole} = 6.2 \text{ liters } H_2 \text{ at STP.}$$

$V_{20^\circ \text{C}, 765 \text{ mm}} = 6.2 \text{ liters} \times 760 \text{ mm}/(765 \text{ mm} - 18 \text{ mm}) \times 293^\circ\text{K}/273^\circ\text{K} = 6.8$ liters.

35.

$$\underset{1 \text{ mole}}{CS_2(l)} + \underset{3 \text{ moles}}{3O_2(g)} \rightarrow CO_2(g) + 2SO_2(g)$$

By inspection 3.00 moles of O_2 is required. This is 3.00 moles $\times$ 22.4 liters/mole = 67.2 liters at STP.

Volume of air = $67.2 \text{ liters} \times \frac{100\% \text{ air}}{21.0\% \text{ } O_2} = 320 \text{ liters air at STP.}$ $V_{29^\circ \text{C}, 744 \text{ mm}} =$

$320 \text{ liters} \times 760 \text{ mm}/744 \text{ mm} \times 302^\circ\text{K}/273^\circ\text{K} = 362 \text{ liters air.}$

36. By inspection there will be 3.00 moles of CO_2—SO_2 mixture, or 67.2 liters of mixture at STP. $V_{29^\circ \text{C}, 744 \text{ mm}} = 67.2 \text{ liters} \times 760 \text{ mm}/744 \text{ mm} \times 302^\circ\text{K}/273^\circ\text{K} = 75.9 \text{ liters.}$

37.

$$\underset{1 \text{ mole}}{MnO_2(s)} + \underset{4 \text{ moles}}{4HCl(aq)} \rightarrow MnCl_2(aq) + 2H_2O(l) + \underset{1 \text{ mole}}{Cl_2(g)}$$

1 mole MnO_2 = 86.9 g 1 mole HCl = 36.5 g 1 mole Cl_2 = 22.4 liters

(*a*) $X = 1.00 \text{ liter } Cl_2 \times \frac{\text{mole}}{22.4 \text{ liters}} \times \frac{1 \text{ mole } MnO_2}{1 \text{ mole } Cl_2} \times 86.9 \text{ g/mole} = 3.88 \text{ g } MnO_2$

(*b*) $Y = 1.00 \text{ liter } Cl_2 \times \frac{\text{mole}}{22.4 \text{ liters}} \times \frac{4 \text{ moles HCl}}{1 \text{ mole } Cl_2} \times 36.5 \text{ g/mole} = 6.52 \text{ g HCl}$

38. $6.52 \text{ g HCl} \times \frac{100\% \text{ HCl soln}}{37.4\% \text{ HCl}} \times \frac{1 \text{ ml}}{1.189 \text{ g}} = 14.7 \text{ ml HCl soln}$

39. $V_{STP.} = 100 \text{ liters} \times 747 \text{ mm}/760 \text{ mm} \times 273^\circ\text{K}/293^\circ\text{K} = 91.5 \text{ liters}$

$$\underset{1 \text{ mole}}{C(s)} + O_2(g) \rightarrow \underset{1 \text{ mole}}{CO_2(g)}$$

1 mole C = 12.0 g 1 mole CO_2 = 22.4 liters

$$X = 91.5 \text{ liters } CO_2 \times \frac{\text{mole}}{22.4 \text{ liters}} \times \frac{1 \text{ mole C}}{1 \text{ mole } CO_2} \times 12.0 \text{ g/mole} = 49.0 \text{ g C}$$

$$49.0 \text{ g C} \times \frac{100\% \text{ charcoal}}{90.0\% \text{ C}} = 54.4 \text{ g charcoal.}$$

40. 5.00 liters × 600 mm/760 mm × 273°K/293°K × 1 mole/22.4 liters × 71.0 g Cl_2/mole = 11.7 g Cl_2

41. $pV = nRT$

$T = pV/nR$

T = [745 mm/760 mm] × 2.50 liters/0.100 mole × [0.0821 liter-atm/mole-°K]

T = 298° K or 25° C

42. $pV = nRT$ and $n = pV/RT$ where n = no. of moles of gas. Thus $n = m/M$ (mass of gas in grams ÷ g-mol wt). Substituting $m/M = pV/RT$ and $m/V = pM/RT$. By definition: $D = m/V$. Substituting: $D = pM/RT$. At STP, p, R, and T are constants. Therefore D is directly proportional to M.

43. $pV = nRT$ where $n = m/M$ (mass of gas in grams ÷ g-mol wt). $pV = mRT/M$. $M = mRT/pV$. p = 740/760 atm = 0.974 atm.

$$M = \frac{1.98 \text{ g} \times 0.0821 \text{ liter-atm/mole-°K} \times 285° \text{ K}}{0.974 \text{ atm} \times 1.07 \text{ liters}}$$

M = 44.4 g/mole mol wt = 44.4

Chapter 11

Answers to Questions pp. 220-222

Group A

1. (*a*) The forces of attraction between the particles of a liquid are strong enough to permit it to have one free surface; thus it has a definite volume. (*b*) The kinetic energy of the particles of a liquid is sufficient to enable single particles or groups of particles to move with respect to one another. (*c*) The particles of a liquid are almost as close together as it is possible for them to be. (*d*) The particles of a liquid may intermingle with the particles of another liquid. (*e*) The kinetic energy of some particles of a liquid is great enough for the particles to escape from the liquid surface into the space above.
2. Finely divided particles of a solid suspended in a liquid and viewed through a microscope are observed to move about in a random manner. The use of lighter or smaller particles and higher temperatures increases the motion.
3. Two opposing processes in the same system proceeding simultaneously at equal rates.
4. No. The vapor can continually escape, so that the process of condensation can never occur as fast as the process of evaporation. Hence, no equilibrium is possible.
5. The rate of evaporation is reduced and the rate of condensation is increased. This causes the quantity of liquid to increase and the quantity of vapor to decrease. Ultimately the rate of condensation will become equal to the rate of evaporation and equilibrium will be restored at a lower equilibrium vapor pressure.
6. The greater the pressure on the water surface, the higher the boiling temperature of the water.

7. (*a*) The highest temperature at which it is possible to liquefy a gas with any amount of pressure. (*b*) The pressure required to liquefy a gas at its critical temperature. (*c*) The volume occupied by one mole of gas under the conditions of critical temperature and critical pressure.
8. Its temperature must be at or below its critical temperature, and its pressure must be raised to or above the vapor pressure of the liquefied gas at this temperature.
9. (*a*) The particles of a solid are held close together in fixed positions by strong forces. (*b*) Same as (*a*). (*c*) Same as (*a*). (*d*) The particles of a solid are in weak vibratory motion back and forth about fixed equilibrium positions. (*e*) The particles of a solid are held in fixed equilibrium positions which, if in a regular pattern, produce crystals.
10. (*a*) Hard, brittle, rather high melting point, good insulator. (*b*) Low melting point, fairly soft, volatile, good insulator. (*c*) Very hard, brittle, rather high melting point, nonconductor. (*d*) High electric conductivity, hardness and melting point extremely variable.
11. Transparent, odorless, almost colorless, tasteless, freezes at 0° C, boils at 100° C.
12. The volume of the steam is about 1700 times the volume of the water at 100°C and one atm pressure.
13. (*a*) $H_2O(s) + 1.44\ kcal \rightarrow H_2O(l)$
(*b*) $H_2O(l) + 9.70\ kcal \rightarrow H_2O(g)$
14. A water molecule is composed of two atoms of hydrogen joined to an atom of oxygen by covalent bonds, with an angle of about 105° between the bonds. Because of the higher electronegativity of the oxygen, the molecule has regions of partial electric charge which make it polar.
15. (*a*) A weak chemical bond between a hydrogen atom in one polar molecule and a very electronegative atom in a second polar molecule. (*b*) They cause the boiling point to be higher than it otherwise would be.
16. (*a*) A compound that does not decompose easily. Water. (*b*) A compound that decomposes easily. Mercury(II) oxide.
17. (*a*) Sodium, potassium, calcium, lithium, magnesium, etc. (*b*) Sodium, potassium, lithium, and calcium will react with cold water; magnesium will react with boiling water.
18. (*a*) An oxide which forms either a basic solution or an acid solution when combined with water. (*b*) A basic anhydride is an ionic oxide of a metal which forms a basic solution with water, while an acid anhydride is a covalent oxide of a nonmetal which forms an acid solution with water. (*c*) A nonmetallic oxide. (*d*) A metallic oxide.
19. This represents the loose attachment of water molecules to the ions of a hydrate.
20. The chemical changes which occur when water is added to an effervescent alkalizing tablet.
21. No. Sodium carbonate decahydrate crystals are efflorescent and lose water of crystallization to air in which the water vapor pressure is lower than the vapor pressure of the crystals.
22. (*a*) Anhydrous calcium chloride is deliquescent and so removes water vapor from the air, finally forming a concentrated solution. (*b*) The calcium chloride may be placed in bags or perforated metal cones and supported above shallow pans. The solution drips into the pans.
23. By electrolysis.
24. As a moderator in nuclear reactors and as a "tracer" in research work in physiology.

Group B

25. (*a*) Low-molecular-weight nonpolar covalent molecules. (*b*) Higher-molecular-weight nonpolar covalent molecules and polar covalent molecules. (*c*) Ions, covalent networks, and metallic ions permeated by electron gas.
26. (*a*) About 92° C. (*b*) About 72° C. (*c*) About 28° C.
27. (*a*) Equilibrium shifts to right to absorb energy. (*b*) Equilibrium shifts to right to occupy larger volume. (*c*) No effect, since system is in closed container.
28. (*a*) Work is done on the gas during compression, and this increases its energy. (*b*) Gas molecules do work during expansion, and this decreases their energy.

29. (*a*) No, this temperature is higher than its critical temperature. (*b*) Yes, this temperature is lower than its critical temperature.
30. (*a*) It increases their potential energy as they do work in becoming particles of vapor. (*b*) It increases their potential energy as they do work in becoming particles of liquid.
31. By observing the crystal structure. Sodium chloride forms cubic crystals. Alum forms octahedral crystals.
32. Camphor forms molecular crystals, the lattice of which is bound by weak dispersion interaction forces. The crystal lattice is easily distorted and molecules can easily acquire enough energy to overcome completely the weak dispersion interaction forces and escape as free molecules.
33. Because ice has a rigid open lattice structure in which the molecules are farther apart than in the liquid.
34. (*a*) Equilibrium shifts to right to absorb energy. (*b*) Equilibrium shifts to left to supply energy. (*c*) Equilibrium will shift slightly to right, since water occupies slightly less volume than an equal mass of ice.
35. When water is warmed, the breaking down of some of the hydrogen bonds enables water molecules to crowd closer together, while the increased energy of the molecules causes them to spread apart. Up to 4° C, the first effect predominates causing an increase in density. Above 4° C, the second effect predominates and the density becomes less.
36. They are strong bonds.
37. Tobacco is hygroscopic; thus in damp weather the leaves are not so brittle.
38. Molecules composed of the isotopes of both hydrogen and oxygen as well as ions formed from each of these isotopes.

Solutions to Problems p. 222

Group A

1. Oxygen. 5.0 ml.
2. Hydrogen. 40.0 ml.
3. (*a*) 75.0 ml.

   (*b*) $\underset{\text{1 mole}}{2H_2(g)} + \underset{\text{37.5 ml}}{O_2(g)} \rightarrow \underset{\text{2 moles}}{2H_2O}$ (X mole)

   1 mole O_2 = 22.4 liters

   $$X = 37.5 \text{ ml } O_2 \times \frac{\text{mole}}{22.4 \text{ liters}} \times \frac{1 \text{ liter}}{1000 \text{ ml}} \times \frac{2 \text{ moles } H_2O}{1 \text{ mole } O_2} = 3.35 \times 10^{-3} \text{ mole } H_2O$$

4. (*a*) Oxygen, 25.0 ml.

   (*b*) $\underset{\text{2 moles}}{2H_2(g)} + \underset{\text{50.0 ml}}{O_2(g)} \rightarrow \underset{\text{2 moles}}{2H_2O}$ (X mmole)

   1 mmole H_2 = 22.4 ml

   $$X = 50.0 \text{ ml } H_2 \times \frac{\text{mmole}}{22.4 \text{ ml}} \times \frac{2 \text{ moles } H_2O}{2 \text{ moles } H_2} = 2.23 \text{ mmoles } H_2O$$

5. $\underset{\text{2 moles}}{2H_2} + \underset{\text{1 mole}}{O_2} \rightarrow \underset{\text{2 moles}}{2H_2O}$ (X (g), Y (g), 15.0 moles)

   1 mole H_2 = 2.02 g 1 mole O_2 = 32.0 g

$$X = 15.0 \text{ moles } H_2O \times \frac{2 \text{ moles } H_2}{2 \text{ moles } H_2O} \times 2.02 \text{ g } H_2/\text{mole} = 30.3 \text{ g } H_2$$

$$Y = 15.0 \text{ moles } H_2O \times \frac{1 \text{ mole } O_2}{2 \text{ moles } H_2O} \times 32.0 \text{ g } O_2/\text{mole} = 24\bar{0} \text{ g } O_2$$

6. (*a*) $\frac{15 \text{ Å}^3}{\text{molecule}} \times \frac{6.02 \times 10^{23} \text{ molecules}}{\text{mole}} \times \left(\frac{10^{-8} \text{ cm}}{1 \text{ Å}}\right)^3 \times \frac{\text{ml}}{\text{cm}^3} = 9.0 \text{ ml/mole}$

(*b*) $\frac{18.0 \text{ g}}{\text{mole}} \times \frac{1 \text{ ml } H_2O}{1 \text{ g } H_2O} = 18.0 \text{ ml/mole}$

(*c*) About one-half the volume of water is empty space.

Group B

7. Formula wt $Na_2CO_3 \cdot 10H_2O = 2(23.0) + 12.0 + 3(16.0) + 10(18.0) = 286.0$
Formula wt $Na_2CO_3 = 2(23.0) + 12.0 + 3(16.0) = 106.0$

$$10\bar{0} \text{ g } Na_2CO_3 \cdot 10H_2O \times \frac{106.0 \text{ } Na_2CO_3}{286.0 \text{ } Na_2CO_3 \cdot 10H_2O} = 37.1 \text{ g } Na_2CO_3$$

8. Formula wt $CoCl_2 \cdot 6H_2O = 58.9 + 2(35.5) + 6(18.0) = 237.9$

Co: $\frac{58.9 \text{ Co}}{237.9 \text{ } CoCl_2 \cdot 6H_2O} \times 100\% \text{ } CoCl_2 \cdot 6H_2O = 24.8\% \text{ Co}$

Cl: $\frac{71.0 \text{ Cl}}{237.9 \text{ } CoCl_2 \cdot 6H_2O} \times 100\% \text{ } CoCl_2 \cdot 6H_2O = 29.8\% \text{ Cl}$

H_2O: $\frac{108.0 \text{ } H_2O}{237.9 \text{ } CoCl_2 \cdot 6H_2O} \times 100\% \text{ } CoCl_2 \cdot 6H_2O = 45.4\% \text{ } H_2O$

9. $ZnSO_4$: $\frac{56.14 \text{ g } ZnSO_4}{161.5 \text{ g/mole}} = 0.348 \text{ mole } ZnSO_4.$

H_2O: $\frac{43.86 \text{ g } H_2O}{18.0 \text{ g/mole}} = 2.44 \text{ moles } H_2O$

$ZnSO_4$: $H_2O = 0.348/0.348{:}2.44/0.348 = 1{:}7.01$
Empirical formula = $ZnSO_4 \cdot 7H_2O$

10. H_2O: $\frac{45.0 \text{ g } H_2O}{124.8 \text{ g } CuSO_4 \cdot 5H_2O} \times 100\% \text{ } CuSO_4 \cdot 5H_2O = 36.1\% \text{ } H_2O$

11. Cu: $\frac{31.8 \text{ g Cu}}{63.5 \text{ g/mole}} = 0.501 \text{ mole Cu}$

S: $\frac{16.0 \text{ g S}}{32.1 \text{ g/mole}} = 0.498 \text{ mole S}$

O: $\frac{32.0 \text{ g O}}{16.0 \text{ g/mole}} = 2.00 \text{ moles O}$

H_2O: $\frac{45.0 \text{ g } H_2O}{18.0 \text{ g/mole}} = 2.50 \text{ moles } H_2O$

$CU:S:O:H_2O = 0.501/0.498:0.498/0.498:2.00/0.498:2.50/0.498 = 1:1:4:5$
Empirical formula = $CuSO_4 \cdot 5H_2O$

12. (*a*) 1 mole CCl_4 = 154.0 g

$$\frac{154.0\text{ g}}{\text{mole}} \times \frac{1}{1.600\text{ g/ml}} \times \frac{1\text{ mole}}{6.023 \times 10^{23}\text{ molecules}} = 1.598 \times 10^{-22}\text{ ml/molecule}$$

(*b*) $V = 4/3\pi r^3 = \pi d^3/6$. $d = \sqrt[3]{6V/\pi}$.

$d = \sqrt[3]{6 \times 1.598 \times 10^{-22}\text{ ml}/3.1416} = 6.73 \times 10^{-8}$ cm = 6.73 Å

13. Radius of atom of Hg = 1.49 Å
(*a*) $V = 4/3\pi r^3$

$$V = \frac{4 \times 3.14 \times (1.49\text{ Å})^3}{3} = 13.9\text{ Å}^3$$

(*b*) Effective volume of Hg atom in cubic array is that of a cube 2×1.49 Å on a side = $(2.98\text{ Å})^3 = 26.5\text{ Å}^3$

$$\frac{6.02 \times 10^{23}\text{ atoms}}{1.00\text{ mole}} \times \frac{26.5\text{ Å}^3}{\text{atom}} \times \frac{1\text{ ml}}{10^{24}\text{ Å}^3} = 16.0\text{ ml/mole.}$$

Chapter 12

Answers to Questions p. 244

Group A

1. Water, alcohol, gasoline, kerosene, carbon tetrachloride, carbon disulfide, chloroform, ether. Others may be given.
2. They are indefinite and give only general information as to concentration. These terms have no quantitative significance.
3. (*a*) Gas in gas, in liquid, in solid; liquid in gas, in liquid, in solid; solid in gas, in liquid, in solid. (*b*) Solid in liquid is the most common.
4. The carbon dioxide is less soluble in water at atmospheric pressure than at the higher pressure within the soda fountain. Therefore most of the carbon dioxide comes out of solution and escapes as a gas when the carbonated water is drawn from the fountain.
5. Solution equilibrium limits the quantity of solute in solution.
6. The term dissolve refers to the solution process in which a solute goes into solution diffusing throughout the solvent. The term melt refers to the change from solid to liquid state.
7. (*a*) The solubility of a gas in a liquid increases with pressure. (*b*) The effect of pressure on the solubility of a solid in a liquid is negligible.
8. (*a*) The solubility of a gas in a liquid decreases with rise in temperature. (*b*) The solubility of solids in liquids generally increases with rise in temperature, however there are exceptions.
9. (*a*) The term miscible refers to the mutual solubility of two liquids. Immiscibility refers to liquids which do not mix. (*b*) Miscible: alcohol and water, alcohol and ether. Immiscible: oil and water, water and carbon tetrachloride.
10. Polar molecules are covalent structures having a negative region and a positive region.
11. Alcohol is a volatile substance; thus the boiling point of the solution will be lower than that of water.
12. Under existing conditions the rate at which a solute goes into solution and the rate at which the solute separates from solution may become equal. When these opposing processes are equal, solution equilibrium exists and the concentration is the greatest possible.

Group B

13. (*a*) The partial pressure of oxygen in the atmosphere determines the amount of oxygen dissolved since they are in equilibrium. (*b*) Oxygen would come out of solution because of the shift in equilibrium.
14. Powder the copper(II) sulfate crystals, add to hot water, and stir the mixture.
15. The carbon tetrachloride molecule is a regular tetrahedral structure in which the four polar bonds are symmetrically arranged.
16. The alcohol molecule has a region which is distinctly nonpolar and a region which is polar. The molecule of carbon tetrachloride is nonpolar and that of water is polar. Applying the rough rule that like dissolves like, the solubility of carbon tetrachloride may be due to the nonpolar region of the alcohol molecule. The solubility of water may be due to the polar region of the alcohol molecule.
17. The solubility of carbon dioxide decreases as the temperature rises. The gas collects in the neck of the bottle and may build up enough pressure to blow off the cap.
18. Calcium hydroxide is more soluble in cold water.
19. The number of grams of the salt that can be dissolved in 100 g of water at a given temperature is plotted as the ordinate against that temperature in $^\circ$ C as the abscissa. A succession of such points is plotted and the points are then joined by a smooth curve.
20. (*a*) The mixture of two different kinds of molecules has greater randomness than the separate pure liquids and is, therefore, more probable. The entropy is increased by formation of the solution. (*b*) The entropy change favors the solution since entropy is increased in this dissolving process. (*c*) The dissolving action is exothermic. The negative sign means that the solution has less energy than the separate components. Thus, heat is given up during dissolving. (*d*) The energy change is toward minimum energy and favors the solution. (*e*) Both the entropy change toward maximum entropy and the energy change toward minimum energy favor the dissolving process. The two driving forces work together to promote miscibility.

Solutions to Problems p. 245

Group A

1. g-mol wt of C_2H_5OH = 46.0 g

$$\frac{0.175 \text{ mole } C_2H_5OH}{\text{kg } H_2O} \times \frac{46.0 \text{ g}}{\text{mole}} \times 40\bar{0} \text{ g } H_2O \times \frac{\text{kg}}{10^3 \text{ g}} = 3.21 \text{ g } C_2H_5OH$$

2. g-mol wt of $C_{12}H_{22}O_{11}$ = 342 g

$$\frac{0.100 \text{ mole } C_{12}H_{22}O_{11}}{\text{kg } H_2O} \times \frac{342 \text{ g}}{\text{mole}} \times 250\bar{0} \text{ g } H_2O \times \frac{\text{kg}}{10^3 \text{ g}} = 85.5 \text{ g } C_{12}H_{22}O_{11}$$

3. g-mol wt of $C_6H_{12}O_6$ = 18$\bar{0}$ g

$$\frac{6.75 \text{ g } C_6H_{12}O_6}{325 \text{ g } H_2O} \times \frac{\text{mole}}{18\bar{0} \text{ g}} \times \frac{10^3 \text{ g}}{\text{kg}} = \frac{0.115 \text{ mole } C_6H_{12}O_6}{\text{kg } H_2O} = 0.115 \ m$$

4. g-mol wt of $C_3H_5(OH)_3$ = 92.0 g

$$\frac{46.0 \text{ g } C_3H_5(OH)_3}{75\bar{0} \text{ g } H_2O} \times \frac{\text{mole}}{92.0 \text{ g}} \times \frac{10^3 \text{ g}}{\text{kg}} = \frac{0.667 \text{ mole } C_3H_5(OH)_3}{\text{kg } H_2O} = 0.667 \ m$$

5. g-mol wt of CH_3OH = 32.0 g

$$\frac{96.0 \text{ g } CH_3OH}{350\bar{0} \text{ g } H_2O} \times \frac{\text{mole}}{32.0 \text{ g}} \times \frac{10^3 \text{ g}}{\text{kg}} = \frac{0.857 \text{ mole } CH_3OH}{\text{kg } H_2O} = 0.857 \ m$$

6. g-mol wt of $C_6H_{12}O_6$ = 180 g

$$\frac{\text{kg } H_2O \times 90.0 \text{ g } C_6H_{12}O_6}{0.250 \text{ mole } C_6H_{12}O_6} \times \frac{\text{mole}}{180 \text{ g}} \times \frac{10^3 \text{ g}}{\text{kg}} = 2000 \text{ g } H_2O$$

7. g-mol wt of $C_{10}H_8$ = 128 g

$$\frac{\text{kg } C_6H_6 \times 32.0 \text{ g } C_{10}H_8}{0.400 \text{ mole } C_{10}H_8} \times \frac{\text{mole}}{128 \text{ g}} \times \frac{10^3 \text{ g}}{\text{kg}} = 625 \text{ g } C_6H_6$$

8. g-mol wt $C_2H_4(OH)_2$ = 62.0 g

$$\frac{31.0 \text{ g } C_2H_4(OH)_2}{100 \text{ g } H_2O} \times \frac{\text{mole}}{62.0 \text{ g}} \times \frac{10^3 \text{ g}}{\text{kg}} = \frac{5.00 \text{ mole } C_2H_4(OH)_2}{\text{kg } H_2O} = 5.00\ m$$

9. g-mol wt of I_2 = 254 g

$$\frac{0.762 \text{ g } I_2}{450 \text{ g } CCl_4} \times \frac{\text{mole}}{254 \text{ g}} \times \frac{10^3 \text{ g}}{\text{kg}} = \frac{0.00667 \text{ mole } I_2}{\text{kg } CCl_4} = 0.00667\ m$$

10. g-mol wt $C_{12}H_{22}O_{11}$ = 342 g

$$\Delta T_f = K_f m = \frac{1.86\ C^\circ}{\text{mole } C_{12}H_{22}O_{11}/\text{kg } H_2O} \times \frac{15.0 \text{ g } C_{12}H_{22}O_{11}}{150.0 \text{ g } H_2O} \times \frac{\text{mole}}{342 \text{ g}} \times \frac{10^3 \text{ g}}{\text{kg}}$$

$\Delta T_f = 0.545\ C^\circ$
f.p. = -0.545° C

11. $\Delta T_b = K_b m = \dfrac{0.51\ C^\circ}{\text{mole } C_{12}H_{22}O_{11}/\text{kg } H_2O} \times \dfrac{15.0 \text{ g } C_{12}H_{22}O_{11}}{150.0 \text{ g } H_2O} \times \dfrac{\text{mole}}{342 \text{ g}} \times \dfrac{10^3 \text{ g}}{\text{kg}}$

12. $\Delta T_f = K_f m = \dfrac{1.86\ C^\circ}{\text{mole solute/kg } H_2O} \times \dfrac{11.25 \text{ g solute}}{250 \text{ g } H_2O} \times \dfrac{\text{mole}}{180 \text{ g}} \times \dfrac{10^3 \text{ g}}{\text{kg}}$

$\Delta T_f = 0.46\ C^\circ$
f.p. = -0.46° C

13. ΔT_f for benzene = 5.5° C $- 4.3^\circ$ C = $1.2\ C^\circ$
$\Delta T_f = K_f m$

$$m = \frac{\Delta T_f}{K_f} = \frac{1.2\ C^\circ}{5.1\ C^\circ/\text{molal}} = 0.24 \text{ molal}$$

14. ΔT_b for water = 100.11° C $- 100.00^\circ$ C = $0.11\ C^\circ$
$\Delta T_b = K_b m$

$$m = \frac{\Delta T_b}{K_b} = \frac{0.11\ C^\circ}{0.51\ C^\circ/\text{molal}} = 0.22 \text{ molal}$$

Group B

15. (*a*) $C:H:O = \frac{32.0}{12.0} : \frac{4.0}{1.0} : \frac{64.0}{16.0} = 2.67:4.0:4.00$

$$= \frac{2.67}{2.67} : \frac{4.0}{2.67} : \frac{4.00}{2.67} = 1.0:1.5:1.5$$

$$= 2:3:3 \text{ empirical formula } = C_2H_3O_3$$

(*b*) $$\text{g-mol wt} = \frac{1.86\ C^\circ \times 15.0 \text{ g solute}}{\text{mole solute/kg } H_2O \times 0.186\ C^\circ \times 1.00 \text{ kg } H_2O}$$

g-mol wt = 150 g/mole
mol wt = 150

(*c*) $(C_2H_3O_3)_x = 150 \quad x = 150 \div 75 = 2.0$
molecular formula = $(C_2H_3O_3)_2$ or $C_4H_6O_6$ ($H_2C_4H_4O_6$, tartaric acid)

16. (*a*) $C:H:O = \frac{40.00}{12.0} : \frac{6.67}{1.0} : \frac{53.33}{16.0} = 3.33:6.7:3.33 = 1:2:1$

empirical formula = CH_2O

(*b*) $$\text{g-mol wt} = \frac{0.51\ C^\circ \times 9.0 \text{ g solute}}{\text{mole solute/kg } H_2O \times 0.051\ C^\circ \times 500 \text{ g } H_2O \times \text{kg}/10^3 \text{ g}}$$

g-mol wt = 180 g/mole
mol wt = 180

(*c*) $(CH_2O)_x = 180 \quad x = 180 \div 30 = 6.0$
molecular formula = $(CH_2O)_6$ or $C_6H_{12}O_6$

17. (*a*) $C:H:Br = \frac{30.4}{12.0} : \frac{1.69}{1.0} : \frac{68}{80} = 2.53:1.7:0.85$

$$= \frac{2.53}{0.85} : \frac{1.7}{0.85} : \frac{0.85}{0.85} = 3:2:1 \text{ empirical formula } C_3H_2Br$$

(*b*) $$\text{g-mol wt} = \frac{5.12\ C^\circ \times 10.0 \text{ g solute}}{\text{mole solute/kg } C_6H_6 \times 2.1\ C^\circ \times 100 \text{ g } C_6H_6 \times \text{kg}/10^3 \text{ g}}$$

g-mol wt = 240 g/mole
mol wt = 240

(*c*) $(C_3H_2Br)_x = 240$
$x = 240 \div 118 = 2.0$
molecular formula = $(C_3H_2Br)_2$ or $C_6H_4Br_2$

Chapter 13

Answers to Questions pp. 258-259

Group A

1. An electrolyte is a substance whose water solution conducts electricity. Solutions of nonelectrolytes do not conduct electricity.
2. Electrolytes lower the freezing point and raise the boiling point of their solvents more than nonelectrolytes.
3. The theory of ionization.
4. (*a*) That electrolytes in solution exist in the form of ions. (*b*) That an ion is an

atom or a group of atoms which carries an electric charge. (*c*) That the water solution of an electrolytes contains an equal number of positive and negative charges.

5. An ion is an atom or a group of atoms which carries an electric charge.
6. $2H_2O \rightarrow H_3O^+ \ OH^-$.
7. The greater positive nuclear charge of the oxygen atom attracts the electrons shared with the two hydrogen atoms (covalent bonds) more strongly. This causes the oxygen end of the molecule to be more negative, and the hydrogen end to be more positive.
8. It is a lattice-like structure of positive and negative ions arranged in a pattern determined by their oxidation states and their size.
9. An atom is electrically neutral, and except for the noble gases, has an unstable outer shell of electrons. An ion is electrically unbalanced, having an excess of either positive or negative charges, and a stable outer shell of electrons.
10. As an atom gains or loses electrons it forms an ion having the stable electron configuration of a noble gas.
11. (*a*) The water dipoles surround the ions of an electrovalent compound and weaken the bonds between them. The hydrated ions diffuse throughout the solvent as the electrovalent compound dissolves. (*b*) The process is reversed by evaporating the solvent.
12. All the dissolved particles exist as ions in the solution.
13. Potassium chloride is an ionic crystalline substance. When fused, its ions are free to move about and conduct an electric current.
14. (*a*) 0.0372 C°, (*b*) 0.0186 C°, (*c*) 0.0930 C°.

Group B

15. (*a*) Water molecules are polar. The more negative oxygen end of several water molecules is attracted to the more positive hydrogen end of a hydrogen chloride molecule. The more positive hydrogen end of several other water molecules is attracted to the more negative chlorine end of a hydrogen chloride molecule. The collection of such dipoles around the hydrogen chloride molecule weakens the attractive force between the hydrogen and chloride, and the bond between them is broken. The chlorine keeps the electron from the hydrogen because of its stronger positive nucleus. Both the hydrogen and chlorine atoms are thus converted to hydrated ions. (*b*) $HCl + H_2O \rightarrow H_3O^+ \ Cl^-$.
16. Dissociation is the separation of ions already existing in the solid by the action of the solvent. Ionization is the formation of ions from covalent molecules by the action of the solvent.
17. A solution equilibrium condition exists in which the rate at which Na^+ ions and NO_3^- ions are being hydrated equals the rate at which these hydrated ions are rejoining the crystal lattice of the undissolved sodium nitrate.

$$Na^+ NO_3^- \rightleftarrows Na^+ + NO_3^-$$

18. (*a*) They are covalent molecules which have a symmetrical bond distribution and thus have a net structure which is nonpolar.
    (*b*) They do not ionize because of the nonpolar nature of the molecule as a whole. They are not likely to be very soluble in polar solvents and such solvent molecules can exert little influence on them.
19. The freezing-point depression and boiling-point elevation are dependent on the number of particles in solution, not on their identity. Electrolytes provide more particles per mole, due to dissociation or ionization, than do nonelectrolytes. Thus electrolytes cause a greater freezing-point depression and boiling-point elevation than do nonelectrolytes.
20. (*a*) $Ca^{++} \ Cl_2^- \rightarrow Ca^{++} + 2Cl^-$. (*b*) The dissociation results in 3 ions, thus the freezing-point depression will be 3×1.86 C° or $-5.58°$ C (ideally).
21. By measuring the freezing-point depression or boiling-point elevation, or by measuring the electric conductivity.

22. It is assumed that the hydrated ions in concentrated solutions interfere with each other's activity, causing the apparent ionization to be less than the actual ionization.
23. A concentrated solution of a weak electrolyte is one in which there is a large amount of solute which does not readily ionize. A dilute solution of a strong electrolyte contains a small amount of solute which readily ionizes.
24. (*a*) The temperature of the solution is lowered.
(*b*) The lattice energy is greater than the hydration energy.
(*c*) The increase in entropy is the driving force for dissolving.
25. (*a*) The solubility of KNO_3 increases with temperature.
(*b*) By Le Chatelier's principle raising the temperature of the KNO_3 solution increases the action which takes up heat, the endothermic dissolving process.
26. The substance is an electrolyte and the solution will be a conductor of electricity. The greater freezing-point depression is due to the presence of ions in solution, giving a greater number of particles per mole than nonelectrolytes. The depression per mole of solute is approximately twice the molal-freezing point constant; the ionic charges have a 1:1 ratio.

Solutions to Problems p. 259

Group B

1. (*a*)

$$\frac{10.0 \text{ ml}}{185 \text{ drops}} \times \frac{1 \text{ g}}{\text{ml}} \times \frac{\text{mole}}{18.0 \text{ g}} \times \frac{6.02 \times 10^{23} \text{ molecules}}{\text{mole}} = 1.81 \times 10^{21} \frac{\text{molecules}}{\text{drop}}$$

(*b*)

$$\frac{1.81 \times 10^{21} \text{ molecules}}{\text{drop}} \times \frac{2 \ H_3O^+ \text{ ion}}{10^9 \text{ molecules}} = \frac{4 \times 10^{12} \ H_3O^+ \text{ ions}}{\text{drop}}$$

(*c*)

4×10^{12} OH^- ions/drop

2. $\frac{0.0155 \text{ mole } CuSO_4 \cdot 5H_2O}{\text{kg } H_2O} \times \frac{249.7 \text{ g}}{\text{mole}} \times 125 \text{ g } H_2O \times \frac{\text{kg}}{10^3 \text{ g}} = 0.483 \text{ g}$ $CuSO_4 \cdot 5H_2O$

3. $\Delta T_f = K_f m = \frac{1.86 \text{ C}^\circ}{\text{mole } C_2H_5OH/\text{kg } H_2O} \times \frac{12.0 \text{ g } C_2H_5OH \times \text{mole}/36.0 \text{ g}}{0.600 \text{ kg } H_2O} =$ 103 C°

fp = -1.03° C

4. $\text{C:H:Cl} = \frac{10.1 \text{ g}}{12.0 \text{ g/mole}} : \frac{0.846 \text{ g}}{1.01 \text{ g/mole}} : \frac{89.1 \text{ g}}{35.5 \text{ g/mole}}$

$= 0.842 \text{ mole} : 0.840 \text{ mole} : 2.52 \text{ mole}$

$= 1 : 1 : 3$

Empirical formula = $CHCl_3$ Empirical formula weight = 119.5

$\Delta T_f = K_f m$ where $m = \frac{\text{g solute/g-mol wt}}{\text{kg solvent}}$

$$\text{g-mol wt} = \frac{K_f \text{ g solute}}{\Delta T_f \text{ kg solvent}} = \frac{5.12 \text{ C}^\circ \times 2.50 \text{ g solute}}{\frac{\text{mole solute}}{\text{kg solvent}} \times 1.07 \text{ C}^\circ \times 0.100 \text{ kg solvent}}$$

g-mol wt = $12\bar{0}$ g/mole
mol wt = $12\bar{0}$ Therefore, molecular formula = $CHCl_3$

5. $V_{STP} = 15.0 \text{ ml HCl} \times \frac{748 \text{ mm}}{760 \text{ mm}} \times \frac{273^\circ \text{ K}}{294^\circ \text{ K}} = 13.7 \text{ ml HCl at STP}$

$D_{HCl} = 1.64$ g/liter
$m = DV = 1.64 \text{ g/liter} \times 13.7 \text{ ml HCl} \times \text{liter}/10^3 \text{ ml} = 0.0225 \text{ g HCl}$
(1 mole HCl → 2 moles solute ions)

$$\Delta T_f = 2K_f m = \frac{2 \times 1.86 \text{ C}^\circ}{\text{mole HCl/kg } H_2O} \times \frac{0.0225 \text{ g HCl} \times \text{mole}/36.5 \text{ g}}{1.00 \text{ kg } H_2O} = 0.00230 \text{ C}^\circ$$

Chapter 14

Answers to Questions pp. 284-285

1. Sulfuric Acid—widely used in industry.
   Nitric Acid—making dyes and explosives.
   Hydrochloric acid—cleaning metals.
2. The hydronium ion.
3. An acid gives up protons to particles capable of accepting them; hence it is a proton donor.
4. (*a*) An acid anhydride is a nonmetallic oxide which reacts with water to form an acid. (*b*) A basic anhydride is a metallic oxide which reacts with water to form the basic OH^- ion. (*c*) Examples of acid anhydrides: SO_2, CO_2, P_2O_5. Examples of basic anhydrides: CaO, MgO.
5. (*a*) Binary acids are named by using the prefix hydro-, the root of the name of the element other than hydrogen in the acid, and the ending -ic. (*b*) Oxyacids are named by a system which depends on the amount of oxygen in the acid. The prefix per-, the root of the name of the third element in the acid, and the ending -ic are used with the acid containing the greatest amount of oxygen. Only the root and ending -ic are used with the acid containing the next smaller amount of oxygen. The root and the ending -ous are used for the next acid. The acid containing the smallest amount of oxygen is named by using the prefix hypo-, the root of the name of the third element, and the ending -ous.
6. Bases are substances which react with acids by acquiring protons; bases are proton acceptors.
7. $H_3O^+ + OH^- \rightarrow 2H_2O$
8. (*a*) $Al(OH)_3$. (*b*) H_3AlO_3. (*c*) Amphiprotic.
9. All true salts are electrovalent compounds composed of positive ions, other than hydronium ions, and negative ions, other than hydroxide ions. They are completely ionized in water solutions and as solids are in the form of ionic crystals.
10. Salts are named by combining the names of the positive and negative ions of which they are composed.
11. Ionic reaction: Sodium sulfate and calcium chloride are soluble salts and are readily available. There is little energy change in the reaction. Calcium sulfate is only very slightly soluble and will precipitate while sodium chloride will remain in solution. Calcium sulfate may be recovered by filtering, washing, and drying.

$$2Na^+ + SO_4^{--} + Ca^{++} + 2Cl^- \rightarrow 2Na^+ + 2Cl^- + CaSO_4(s)$$

12. No. $BaSO_4$ is insoluble.

Group B

13. HCl consists of polar molecules which are ionized in water solution due to the hydrating action of the water dipoles. The water solution contains hydrated protons giving it acidic properties.
14. (*a*) When the HCl dissolves in ammonia it gives up a proton to ammonia, forming the ammonium ion and a chloride ion. The HCl acts as a proton donor and in this sense is an acid. (*b*) $HCl + NH_3 \rightarrow NH_4^+ + Cl^-$.

15. (*a*) A water molecule may yield a proton to an ammonia molecule dissolved in it to form the ammonium ion and the hydroxide ion. Since the water molecule donates a proton, it may be called an acid.

(*b*) $H{:}\ddot{N}{:}H + H{:}\ddot{O}{:} \rightarrow H{:}\ddot{N}{:}H^+ + :\ddot{O}{:}^-$ (H above N; H, H, H below N; H below O; H below O)

16. (*a*) Greasy sink traps are cleaned with sodium hydroxide. The hydroxide reacts with the grease, forming soap, which may be washed down the drain. (*b*) Ammonia water may be used to remove grease spots from clothing. It is weakly basic and leaves no residue in the fabric on evaporation.
17. Calcium hydroxide solution or ammonia water may be used to neutralize acid stains on clothing. Both are relatively harmless to fabrics.
18. NaCl—soluble. $CaCO_3$—partly soluble. $BaSO_4$—insoluble. $(NH_4)_2S$—soluble. $Al(C_2H_3O_2)_3$—soluble. Ag_2SO_4—partly soluble. $Pb(NO_3)_2$—soluble. Hg_2Cl_2—insoluble. $Mg_3(PO_4)_2$—partly soluble. CuS—insoluble.
19. (*a*) Hydrogen selenide (hydroselenic acid). (*b*) Iodic acid. (*c*) Gallium hydroxide. (*d*) Cesium hydroxide. (*e*) Radium bromide.
20. (*a*) The H_3O^+ ion. (*b*) The OH^- ion.
21. (*a*) The NH_4^+ ion. (*b*) The NH_2^- ion.
22. $H_2SO_4 + H_2O \rightarrow H_3O^+ + HSO_4^-$
    acid base acid base
    $H_3O^+ + SO_4^{-\ -} \rightarrow HSO_4^- + H_2O$
    acid base acid base

Solutions to Problems p. 285

Group A

1. $\underset{1\text{ mole}}{H_2SO_4} + NaNO_3 \rightarrow \underset{1\text{ mole}}{HNO_3} + \underset{1\text{ mole}}{NaHSO_4}$

   X 50.5 g Y

   $\frac{98.1\text{ g}}{\text{mole}}$ $\frac{63.0\text{ g}}{\text{mole}}$ $\frac{120\text{ g}}{\text{mole}}$

   (*a*) $X = 50.0\text{ g } HNO_3 \times \frac{\text{mole}}{63.0\text{ g}} \times \frac{1\text{ mole } H_2SO_4}{\text{mole } HNO_3} \times \frac{98.0\text{ g}}{\text{mole}} = 77.9\text{ g } H_2SO_4$

   (*b*) $Y = 50.0\text{ g } HNO_3 \times \frac{\text{mole}}{63.0\text{ g}} \times \frac{1\text{ mole } NaHSO_4}{\text{mole } HNO_3} \times \frac{120\text{ g}}{\text{mole}} = 95.2\text{ g } NaHSO_4$

2. $\underset{1\text{ mole}}{25.0\text{ g } CaCO_3} + 2HCl \rightarrow CaCl_2 + H_2O + \underset{1\text{ mole}}{X\ CO_2}$

   $\frac{100\text{ g}}{\text{mole}}$ $\frac{22.4\text{ liters}}{\text{mole}}$

   $X = 25.0\text{ g } CaCO_3 \times \frac{\text{mole}}{100\text{ g}} \times \frac{1\text{ mole } CO_2}{\text{mole } CaCO_3} \times \frac{22.4\text{ liters}}{\text{mole}} = 5.60\text{ liters } CO_2 \text{ at STP}$

   $V' = 56.0\text{ liters} \times \frac{293^\circ\text{ K}}{273^\circ\text{ K}} \times \frac{760\text{ mm}}{745\text{ mm}} = 6.13\text{ liters at } 20^\circ\text{ C and } 745\text{ mm pressure}$

3. $\underset{1\text{ mole}}{\text{CaO}} + \underset{1\text{ mole}}{\text{SiO}_2} \rightarrow \underset{1\text{ mole}}{\text{CaSiO}_3}$

$$\frac{56.1\text{ g}}{\text{mole}} \quad \frac{60.1\text{ g}}{\text{mole}} \quad \frac{116\text{ g}}{\text{mole}}$$

(SiO_2 is in excess)

$$X = 75.0\text{ g CaO} \times \frac{\text{mole}}{56.1\text{ g}} \times \frac{1\text{ mole CaSiO}_3}{\text{mole CaO}} \times \frac{116\text{ g}}{\text{mole}} = 155\text{ g CaSiO}_3$$

4. (*a*) $V_{STP} = 75.0\text{ liters} \times 755\text{ mm}/760\text{ mm} \times 273^\circ\text{ K}/298^\circ\text{ K}$

$V_{STP} = 68.2$ liters

$$\underset{1\text{ mole}}{\text{X CaO}} + \underset{1\text{ mole}}{\text{68.2 liters CO}_2} \rightarrow \underset{1\text{ mole}}{\text{Y CaCO}_3}$$

$$\frac{56.1\text{ g}}{\text{mole}} \quad \frac{22.4\text{ liters}}{\text{mole}} \quad \frac{100\text{ g}}{\text{mole}}$$

$$X = 68.2\text{ liters} \times \frac{\text{mole}}{22.4\text{ liters}} \times \frac{1\text{ mole CaO}}{\text{mole CO}_2} \times \frac{56.1\text{ g}}{\text{mole}} = 171\text{ g CaO}$$

(*b*)
$$Y = 68.2\text{ liters} \times \frac{\text{mole}}{22.4\text{ liters}} \times \frac{1\text{ mole CaCO}_3}{\text{mole CO}_2} \times \frac{100\text{ g}}{\text{mole}} = 304\text{ g CaCO}_3$$

Chapter 15

Answers to Questions pp. 308-309

Group A

1. Molality: moles of solute per kilogram of solvent.
   Molarity: moles of solute per liter of solution.
2. (*a*) One equivalent of an acid is the number of grams required to supply 1 mole of H_3O^+ ions. (*b*) One equivalent of a base is the number of grams required to accept one mole of protons, or to supply one mole of OH^- ions.
3. (*a*) 82.0 g (*b*) 32.7 g (*c*) 61.0 g (*d*) 74.6 g (*e*) 6.9 g (*f*) 19.0 g (*g*) 15.0 g (*h*) 17.0 g (*i*) 60.0 g (*j*) 57.0 g.
4. (*a*) 2 (*b*) 1 (*c*) 2 (*d*) 3 (*e*) 2 (*f*) 1 (*g*) 1 (*h*) 2 (*i*) 2 (*j*) 3.
5. $H_2SO_4 + H_2O \rightarrow H_3O^+ + HSO_4^-$
   $HSO_4^- + H_2O \rightarrow H_3O^+ + SO_4^{--}$
6. CO_3^{--} ion.
7. Ammonium ion concentration in moles per liter.
8. (*a*) pH is the symbol for the hydronium ion index. It is the common logarithm of the reciprocal of the hydronium ion concentration of a solution. $pH = \log \frac{1}{[H_3O^+]}$, where $[H_3O^+]$ is expressed in gram-ion per liter. (*b*) The usual range of the pH scale is from 0 to 14.
9. (*a*) pH = 3. (*b*) $[OH^-] = 1 \times 10^{-11}$ mole/liter.
10. (*a*) The transition interval lies between pH of 3.0 and 4.4. The pH at the equivalence point in this titration is above 7.
    (*b*) Phenol red or phenolphthalein.

Group B

11. The strength of an acid depends on the extent to which ionization occurs in water solution. HCl is highly ionized, yielding a low concentration of hydronium ions.
12. (*a*) Place several pieces of red and blue litmus paper in the bottom of a beaker, and cover with a piece of moist filter paper. In another beaker mix several teaspoons of soil with enough water to make a thin paste. Pour this mixture onto the filter paper in the first beaker. Allow to stand for 20 minutes, and observe the litmus paper through the bottom of the beaker. If it has changed to red, the soil is acid; if to blue, the soil is basic. (*b*) Slaked lime or ground limestone may be added to the soil if acidic.
13. Student answer. Toothpastes are likely to be basic.
14. (*a*) The K^+ ion and the Cl^- ion take no part in the neutralization reaction. They are present as ions in the water solution when the reaction is complete. (*b*) KCl may be recovered as a crystalline salt by evaporation of the water after the neutralization reaction is complete.
15. Litmus. A neutralization reaction between a strong acid and a strong hydroxide has an equivalence point at a pH of about 7. The color change for litmus occurs at pH of 6-8, thus the color change would show that equivalent quantities of the two solutes were present. Bromthymol blue or phenol red may be selected for the same reason.
16. (*a*) 1 mole of NaOH. (*b*) 2 moles of NaOH. (*c*) 3 moles of NaOH.
    (*d*) $NaOH + HCl \rightarrow NaCl + H_2O$
    $2NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2H_2O$
    $3NaOH + H_3PO_4 \rightarrow Na_3PO_4 + 3H_2O$
17. (*a*) Formula:$Ca(OH)_2$. Formula weight = 74. One equiv = 74 g ÷ 2 = 37 g. 0.01-*N* solution requires 0.01 · 37 g/liter = 0.37 g of calcium hydroxide per liter of solution. (*b*) 0.74 g per liter of solution.
18. (*a*) Ten million, or 10^7, liters of water, (*b*) 1 equiv of hydronium. (*c*) 19 g of H_3O^+ ion. (*d*) 1 mole of OH^- ion. (*e*) 1 equiv of hydroxide. (*f*) 17 g of OH^- ion.
19. (*a*) 0.0040-*M* H_3PO_4 is 0.012 *N*. (*b*) 0.15-*M* KOH is 0.15 *N*. (*c*) 2-*M* H_2SO_4 is 4*N*.
20. (*a*) 0.002-*M* H_3PO_4. (*b*) 0.0006-*M* $Al_2(SO_4)_3$. (*c*) 0.015-*M* $Ba(OH)_2$.

Solutions to Problems, pp. 309-310

Group A

1. X 54.75 g
   $NaOH + HCl \rightarrow NaCl + H_2O$
   1 mole 1 mole

   1 mole NaOH = 40.0 g 1 mole HCl = 36.5 g

   (*a*) $X = 54.75 \text{ g HCl} \times \frac{\text{mole}}{36.5 \text{ g}} \times \frac{1 \text{ mole NaOH}}{\text{mole HCl}} \times \frac{40.0 \text{ g}}{\text{mole}} = 60.0 \text{ g NaOH}$

   (*b*) $54.75 \text{ g HCl} \times \frac{\text{mole}}{36.5 \text{ g}} = 1.50 \text{ moles HCl} = 1.50 \text{ moles NaOH}$

2. Formula: KNO_3. Formula weight = 101
   0.250 mole = 0.250 × 101 g = 25.2 g KNO_3/kg H_2O
   or 12.6 g KNO_3/500 g H_2O
3. 1 mole of $C_{12}H_{22}O_{11}$ = 342 g

   $0.800 \times \frac{342 \text{ g/liter}}{1000 \text{ ml/liter}} \times 50.0 \text{ ml} = 13.7 \text{ g sugar.}$

4. $M = \frac{49.0 \text{ g}}{98.1 \text{ g/mole} \times 3.00 \text{ liters}} = 0.167 \frac{\text{mole}}{\text{liter}} = 0.167 \text{ molar}$

5. $M = \dfrac{446\text{ g}}{223\text{ g/mole} \times 5.00\text{ liters}} = 0.400\ \dfrac{\text{mole}}{\text{liter}} = 0.400$ molar

6. $X = 0.500\ \dfrac{\text{mole}}{\text{liter}} \times 0.250\text{ liter} \times \dfrac{58.5\text{ g}}{\text{mole}} = 7.32\text{ g}$

7. $X = 0.300\ \dfrac{\text{mole}}{\text{liter}} \times 0.800\text{ liter} \times \dfrac{666\text{ g}}{\text{mole}} = 160\text{ g}$

8. (*a*) 39.1 g; (*b*) 20.0 g; (*c*) 58.5 g; (*d*) 79.8 g; (*e*) 143 g; (*f*) 90.1 g

9. $N = \dfrac{4.0\text{ g}}{71\text{ g/equiv} \times 1.0\text{ liter}} = 0.056$ equiv/liter = 0.056 normal

10. $N = \dfrac{710\text{g}}{71.0\text{g/equiv} \times 15.0\text{ liters}} = 0.667$ equiv/liter = 0.667 normal

11. $X = 0.100\ \dfrac{\text{equiv}}{\text{liter}} \times 0.500\text{ liter} \times \dfrac{125\text{ g}}{\text{equiv}} = 6.25\text{ g}$

12. $X = 0.500\ \dfrac{\text{equiv}}{\text{liter}} \times 0.200\text{ liter} \times \dfrac{90.1\text{ g}}{\text{equiv}} = 9.01\text{ g}$

13. (*a*) Assuming complete ionization of a dilute solution of a strong acid, $[H_3O^+]$ = 0.01 mole/liter = 10^{-2} mole/liter

$$pH = \log \frac{1}{[H_3O^+]} = \log \frac{1}{10^{-2}} = \log 10^2$$

pH = 2

(*b*) NaOH is an ionic compound, all water solutions are completely ionized. $[OH^-]$ = 0.01 mole/liter = 10^{-2} mol/liter

(*c*) $[H_3O^+]\ [OH^-] = 10^{-14}$ (constant)

$[H_3O^+] = 10^{-14} \div 10^{-2} = 10^{-12}$

$$pH = \log \frac{1}{[H_3O^+]} = \log \frac{1}{10^{-12}} = \log 10^{12}$$

pH = 12

14. $30.0\text{ ml} \times \dfrac{0.500\ N}{0.150\ N} = 100$ ml of 0.150-*N* hydroxide solution required.

15. $0.50\ N \times \dfrac{34\text{ ml}}{100\text{ ml}} = 1.7$-*N* ammonia-water solution.

Group B

16. 1 liter of a 1-*M* solution of any molecular solute contains 6.02×10^{23} molecules of solute (the Avogadro Number). 1 ml of 0.1-*M* solution contains $0.0001 \times 6.02 \times 10^{23}$ molecules = $10^{-4} \times 6.02 \times 10^{23} = 6.02 \times 10^{19}$ molecules of solute per milliliter.

17. Pressure of the dry gas = 745.0 mm − 17.5 mm = 727.5 mm.

$$\text{Vol. of } H_2 \text{ at STP} = 2.55\text{ liter} \times \frac{727.5\text{ mm}}{760\text{ mm}} \times \frac{273^\circ\text{ K}}{293^\circ\text{ K}} = 2.27\text{ liter}$$

$$\underset{2\text{ moles}}{Zn + 2HCl} \rightarrow \underset{1\text{ mole}}{ZnCl_2 + H_2} \quad \overset{2.27\text{ liter}}{\overset{X}{}}$$

1 mole HCl = 36.5 g 1 mole H_2 = 22.4 liters

$$X = 2.27\text{ liters } H_2 \times \frac{\text{mole}}{22.4\text{ liters}} \times \frac{2\text{ moles HCl}}{\text{mole } H_2} \times \frac{36.5\text{ g}}{\text{mole}} = 7.40\text{ g HCl}$$

7.40 g HCl are contained in 400 ml solution

$$\frac{7.40\text{ g HCl}}{400\text{ ml soln}} \times \frac{1000\text{ ml}}{\text{liter}} = 18.5\text{ g HCl/liter solution}$$

$$\frac{18.5\text{ g HCl/liter soln}}{36.5\text{ g/mole}} = 0.507\text{ mole HCl/liter soln} = 0.507\text{-}M\text{ solution}$$

18. (*a*) $\frac{1.00\text{ mole HCl}}{\text{liter}} \times \frac{36.5\text{ g}}{\text{mole}} \times 1.00\text{ liter} = 36.5\text{ g HCl required}$

$$\frac{1.19\text{ g}}{\text{ml}} \times \frac{37.2\text{ g HCl}}{100\text{ g conc soln}} = 0.443\text{ g HCl/ml conc soln}$$

$$36.5\text{ g HCl} \times \frac{\text{ml conc soln}}{0.443\text{ g HCl}} = 82.5\text{ ml conc soln}$$

(*b*) $\frac{3.00\text{ mole HCl}}{\text{liter}} \times \frac{36.5\text{ g}}{\text{mole}} \times 2.00\text{ liters} \times \frac{\text{ml conc soln}}{0.443\text{ g HCl}} = 495\text{ ml conc soln}$

(*c*) $\frac{0.100\text{ equiv HCl}}{\text{liter}} \times \frac{36.5\text{ g}}{\text{g-eq}} \times 5.00\text{ liters} \times \frac{\text{ml conc soln}}{0.443\text{ g HCl}} = 41.2\text{ ml conc soln}$

(*d*) $\frac{0.200\text{ equiv HCl}}{\text{liter}} \times \frac{36.5\text{ g}}{\text{mole}} \times 0.250\text{ liter} \times \frac{\text{ml conc soln}}{0.443\text{ g HCl}} = 0.412\text{ ml conc soln}$

19. (*a*) $\frac{0.500\text{ mole } H_2SO_4}{\text{liter}} \times \frac{98.1\text{ g}}{\text{mole}} \times 2.50\text{ liters} = 123\text{ g } H_2SO_4\text{ required}$

$$\frac{1.84\text{ g}}{\text{ml}} \times \frac{95.0\text{ g } H_2SO_4}{100\text{ g conc soln}} = 1.75\text{ g } H_2SO_4\text{/ml conc soln}$$

$$123\text{ g } H_2SO_4 \times \frac{\text{ml conc soln}}{1.75\text{ g } H_2SO_4} = 70.3\text{ ml conc soln}$$

(*b*) $\frac{0.250\text{ mole } H_2SO_4}{\text{liter}} \times \frac{98.1\text{ g}}{\text{mole}} \times 0.100\text{ liter} \times \frac{\text{ml conc soln}}{1.75\text{ g } H_2SO_4} = 1.40\text{ ml conc soln}$

(*c*) $\frac{1.00\text{ equiv } H_2SO_4}{\text{liter}} \times \frac{49.0\text{ g}}{\text{g-eq}} \times 1.00\text{ liter} \times \frac{\text{ml conc soln}}{1.75\text{ g } H_2SO_4} = 28.0\text{ ml conc soln}$

(*d*) $\frac{0.200\text{ equiv } H_2SO_4}{\text{liter}} \times \frac{49.0\text{ g}}{\text{g-eq}} \times 3.00\text{ liters} \times \frac{\text{ml conc soln}}{1.75\text{ g } H_2SO_4} = 16.8\text{ ml conc soln}$

20. $H_2SO_4 \times 2NaOH \rightarrow Na_2SO_4 + 2H_2O$

$$0.0150 \text{ liter} \times \frac{0.275 \text{ mole } H_2SO_4}{\text{liter}} = 0.00412 \text{ mole } H_2SO_4 \text{ used}$$

2 moles NaOH used per mole H_2SO_4
0.00824 mole NaOH in 20.0 ml NaOH soln
0.00824 mole NaOH/0.0200 liter = 0.412 mole NaOH/liter = 0.412-*M* NaOH

21. Reaction requires 1 mole NaOH per mole $HC_2H_3O_2$.
1 mole $HC_2H_3O_2$ = 60.0 g

$$25.0 \text{ ml NaOH} \times \frac{\text{liter}}{1000 \text{ ml}} \times \frac{0.100 \text{ mole}}{\text{liter}} = 0.00250 \text{ mole NaOH used.}$$

$HC_2H_3O_2$ neutralized = 0.00250 mole
contained in 30.0 ml dilute solution, or 3.00 ml vinegar.

$$\frac{0.00250 \text{ mole } HC_2H_3O_2}{3.00 \text{ ml vinegar}} \times \frac{1000 \text{ ml}}{\text{liter}} = 0.833 \text{ mole } HC_2H_3O_2\text{/liter vinegar}$$

$$\frac{0.833 \text{ mole } HC_2H_3O_2}{\text{liter vinegar}} \times \frac{60.0 \text{ g}}{\text{mole}} = 50.0 \text{ g } HC_2H_3O_2\text{/liter vinegar}$$

Assuming mass of liter of vinegar to be 1000 g;

$$\frac{50.0 \text{ g } HC_2H_3O_2}{1000 \text{ g}} \times 100\% = 5.00\% \ HC_2H_3O_2$$

22. (*a*) $pH = -\log [H_3O^+]$
$[H_3O^+]$ = antilog (−pH) = antilog (−2.9)
$[H_3O^+]$ = antilog (0.1 − 3) = antilog (0.1) × antilog (−3)
$[H_3O^+] = 1.26 \times 10^{-3}$
(*b*) $[H_3O^+]\,[OH^-] = 1.00 \times 10^{-14}$

$$[OH^-] = \frac{1.00 \times 10^{-14}}{1.26 \times 10^{-3}} = 0.795 \times 10^{-11} = 7.95 \times 10^{-12}$$

23. $$[H_3O^+] = \frac{1 \times 10^{-14}}{[OH^-]} = \frac{1 \times 10^{-14}}{2 \times 10^{-2}} = 0.5 \times 10^{-12}$$

$pH = -\log [H_3O^+] = -\log (0.5 \times 10^{-12})$
$pH = -(\log 0.5 + \log 10^{-12}) = -\log \frac{1}{2} - \log 10^{-12}$
$pH = \log 2 + \log 10^{12} = 0.30 + 12 = 12.3$

24. Dilute solutions of strong acids are completely ionized, thus $[H_3O^+] = 0.054$.

$$pH = \log \frac{1}{[H_3O^+]} = -\log [H_3O^+] = -\log 0.054$$

$pH = -(8.73 - 10) = -(-1.27) = 1.27$

25. 25.0 ml of 0.150-*M* NaOH = 0.0250 liter × 0.150 mole/liter
= 0.00375 mole of NaOH.
50.0 ml of 0.100-*M* HCl = 0.0500 liter × 0.100 mole/liter
= 0.00500 mole of HCl.
HCl in excess = 0.00500 mole − 0.00375 mole = 0.00125 mole.

$$\text{Concentration of HCl in resulting solution} = \frac{0.00125 \text{ mole}}{0.0250 \text{ liter} + 0.050 \text{ liter}} =$$

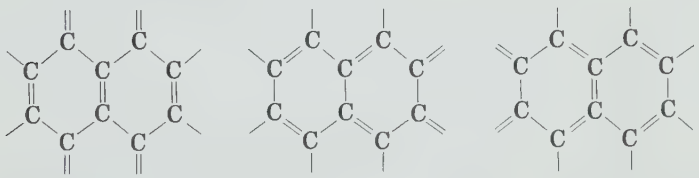
$$\frac{0.00125 \text{ mole}}{0.0750 \text{ liter}} = 0.0166\text{-}M \text{ HCl}$$

$$\text{pH} = -\log(0.0166) = -(8.22 - 10) = (-1.78) = 1.78$$

Chapter 16

Answers to Questions pp. 329-331

Group A

1. (*a*) Because of the large number of compounds of carbon. (*b*) Organic chemistry.
2. They are directed in space toward the four vertices of a regular tetrahedron, if we assume the center of the atom is at the center of the tetrahedron.
3. They join readily with other elements. They also link together in chains, rings, plates, and networks.
4. (*a*) Allotropy is the existence of an element in two or more forms in the same physical phase. (*b*) Diamond and graphite. (*c*) Because X-ray scattering shows that amorphous carbon contains regions in which the carbon atoms are arranged in an orderly way.
5. Because they are one of the hardest materials.
6. 

   It is possible to write three valence bond structures for graphite. Evidence indicates, however, that all the bonds in graphite are the same. None of these structures has an independent existence. The actual structure is a resonance hybrid of the written structures. The carbon-carbon bonds are intermediate in character between single and double bonds.
7. The flat scale-like plates slide over one another easily. It is not soluble or subject to corrosion. It is not affected by heat.
8. (*a*) The process of decomposing materials by heating them in a closed container without access to air or oxygen. (*b*) It is destructive only in the sense that it breaks up or decomposes substances. It is not actually destructive, however, since many new, useful substances are formed by this process.
9. Because the carbon produced by the destructive distillation of coal or wood is not volatile at the temperature at which the destructive distillation is carried out.
10. Adsorption is the concentration of a gas, liquid, or solid on the surface of a liquid or solid with which it is in contact.
11. It usually contains considerable amounts of calcium phosphate.
12. The tremendous internal surface area created during the preparation of activated carbon.

13. The two steps are destructive distillation of a carbon-containing material and treatment of the residue with steam or carbon dioxide at about 100° C.
14. It is used as a filler and reinforcing agent in natural and synthetic rubber.
15. It is used for electrodes in the cells in which aluminum is produced by electrolysis.
16. Because it is a part of the oxygen-carbon dioxide cycle upon which many forms of life depend.
17. (*a*) By burning carbonaceous material, by reaction of steam and natural gas, by fermentation of molasses, and by heating a carbonate. (*b*) By the action of an acid on a carbonate. (*c*) C (combined) + $O_2 \rightarrow CO_2$. $CH_4 + 2H_2O \rightarrow 4H_2 + CO_2$. $C_6H_{12}O_6 \rightarrow 2C_2H_5OH + 2CO_2$. $CaCO_3 \rightarrow CaO + CO_2$. $CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2O + CO_2$.
18. It is a catalyst.
19. (*a*) Carbon dioxide is appreciably soluble in water. (*b*) It will be contaminated with air.
20. Stable gas which neither burns nor supports combustion (except of burning magnesium), anhydride of carbonic acid, forms precipitate of calcium carbonate with limewater, can cause death by suffocation.
21. (*a*) By the reaction of carbon dioxide with water. (*b*) It is a weak acid because the reaction of carbon dioxide with water is slight even though H_2CO_3 ionizes extensively.
22. The reaction of limewater and carbon dioxide to produce a white precipitate.
23. Photosynthesis and the various natural and artificial methods of producing atmospheric carbon dioxide.
24. By blanketing the fire with a dense gas to shut off oxygen. By cooling the burning material below its kindling temperature.
25. (*a*) $2NaHCO_3 + H_2SO_4 \rightarrow Na_2SO_4 + 2H_2O + 2CO_2$. (*b*) Ionic: $2Na^+ + 2HCO_3^- + 2H_3O^+ + SO_4^{--} \rightarrow 2Na^+ + SO_4^{--} + 4H_2O + 2CO_2$. Net ionic: $2HCO_3^- + 2H_3O^+ \rightarrow 4H_2O + 2CO_2$.
26. (*a*) Sodium hydrogen carbonate. (*b*) By reaction with an acid or acid-forming substance in water solution.
27. Most carbon monoxide in the atmosphere comes from automobile engine exhaust. Other sources are other poorly burned fuels, decaying plants, and freshwater algae.
28. The sulfuric acid acts as a dehydrating agent.
29. Hydrogen, carbon monoxide, oxygen, carbon dioxide.
30. As a reducing agent, as a fuel, and for synthesizing organic compounds.

Group B

31. Because of their very strong tendency to share electrons and form covalent bonds.
32. The 2*s* orbital and the three 2*p* orbitals are combined to form four tetrahedrally-oriented equivalent orbitals.
33. (*a*) An element has two or more kinds of molecules, each with different numbers of atoms, or an element has two or more different arrangements of atoms or molecules in a crystal. (*b*) The second reason.
34. When diamonds are burned in pure oxygen, an amount of carbon dioxide is formed which equals the theoretical amount that would be formed if a diamond were pure carbon.
35. In diamond the carbon atoms are covalently bonded to each other in a strong compact fashion. The rigidity of the structure gives diamond its hardness. Since all the electrons are firmly bound in the giant molecular structure, diamond is a nonconductor of electricity. In graphite the carbon atoms are arranged in layers of thin hexagonal plates. The layered structure accounts for the softness. The mobile electrons in the layers account for the electric conductivity. The atoms are closer, on the average, in diamond than in graphite, giving diamond the higher density. The bond strengths in both diamond and graphite are high, giving them both ability to withstand high temperatures without melting.
36. (*a*) Coating the molds used in metal casting, increasing the carbon content of steel, and making clay-graphite crucibles. (*b*) Its high melting point.
37. (*a*) Electrodes for electric-arc steelmaking furnaces. (*b*) Its electric conductivity.
38. The coke.

39. Charcoal will burn. Copper(II) oxide is reduced to copper by hydrogen when it is hot. Manganese dioxide responds to neither of these processes.
40. Activated carbon might be used to recover dry cleaning solvent vapors from exhaust gases. It might be used to purify the dry cleaning solvent.
41. (*a*) A spray of liquid fuel or gaseous fuel vapor, an auxiliary gaseous fuel, and air. (*b*) The fuels supply oxidizable material and the carbon which becomes carbon black. The air is a source of oxygen.
42. The carbon in printer's ink is not soluble.
43. $Ca^{++} CO_3^{--} + 2H_3O^+ \rightarrow Ca^{++} + 3H_2O + CO_2$.
44. $CO_3^{--} + 2H_3O^+ \rightarrow 3H_2O + CO_2$.
45. Its vapor pressure.
46. It is hot enough to decompose carbon dioxide and produce oxygen in which it continues to burn.
47. $CO_2 + H_2O \rightleftarrows H_2CO_3$
$H_2CO_3 + 2H_2O \rightleftarrows 2H_3O^+ + CO_3^{--}$
$2Na^+ OH^- \rightarrow 2Na^+ + 2OH^-$
$2H_3O^+ + 2OH^- \rightarrow 4H_2O$
48. $Ca^{++}(OH^-)_2 \rightarrow Ca^{++} + 2OH^-$
$CO_2 + H_2O \rightleftarrows H_2CO_3$
$H_2CO_3 + 2H_2O \rightleftarrows 2H_3O^+ + CO_3^{--}$
$Ca^{++} + 2OH^- + 2H_2O^+ + CO_3^{--} \rightarrow Ca^{++}CO_3^{--}(s) + 4H_2O$
49. The ring is calcium carbonate formed by the reaction between the limewater and the carbon dioxide in the air. Dilute hydrochloric acid will remove it.
50. Baking soda is sodium hydrogen carbonate. Baking powder is a dry mixture of sodium hydrogen carbonate, a powdered acid or acid-forming substance, and cornstarch.
51. While the carbon-oxygen bonds in carbon dioxide are slightly polar, the linear arrangement of the molecule prevents it from being polar as a whole. The carbon monoxide molecule is unsymmetrical. The 50% contribution of the $^-{:}C{:::}O{:}^+$ structure causes the carbon atom to be more electronegative.
52. Since carbon dioxide molecules are nonpolar and carbon monoxide molecules are only weakly polar, the attractive forces between these molecules are weak, and at rather low temperatures the kinetic energies of the molecules are sufficient to overcome them. Water molecules, on the other hand, are so strongly polar that they form chains by hydrogen bonding. A higher temperature is necessary before the kinetic energy of the molecules can overcome these bonding forces.
53. Carbon dioxide asphyxiates because of lack of oxygen to breathe. Carbon monoxide combines with the hemoglobin of the blood in preference to oxygen, and prevents the hemoglobin from transporting oxygen.

Answers to Problems p. 331

Group A

1. Formula weight of HCOOH = 2(1.0) + 12.0 + 2(16.0) = 46.0

H: $\dfrac{2.0\text{ H}}{46.0\text{ HCOOH}} \times 100\%\text{ HCOOH} = 4.3\%\text{ H}$

C: $\dfrac{12.0\text{ C}}{46.0\text{ HCOOH}} \times 100\%\text{ HCOOH} = 26.1\%\text{ C}$

O: $\dfrac{32.0\text{ O}}{46.0\text{ HCOOH}} \times 100\%\text{ HCOOH} = 69.6\%\text{ O}$

2. (*a*)

X (moles)		2.00 moles		Y (moles)		Z (moles)
Fe_2O_3	+	3CO	→	2Fe	+	$3CO_2$
1 mole		3 moles		2 moles		3 moles

$$X = 2.00 \text{ moles CO} \times \frac{1 \text{ mole } Fe_2O_3}{3 \text{ moles CO}} = 0.667 \text{ mole } Fe_2O_3$$

(*b*) $$Y = 2.00 \text{ moles CO} \times \frac{2 \text{ moles Fe}}{3 \text{ moles CO}} = 1.33 \text{ moles Fe}$$

(*c*) $$Z = 2.00 \text{ moles CO} \times \frac{3 \text{ moles } CO_2}{3 \text{ moles CO}} = 2.00 \text{ moles } CO_2$$

3. $$\underset{1 \text{ mole}}{\overset{12.2 \text{ g}}{ZnO}} + \underset{1 \text{ mole}}{\overset{X}{CO}} \rightarrow \underset{1 \text{ mole}}{\overset{Y}{Zn}} + \underset{1 \text{ mole}}{\overset{Z}{CO_2(g)}}$$

1 mole ZnO = 81.4 g 1 mole CO = 28.0 g

$$X = 12.2 \text{ g ZnO} \times \frac{\text{mole}}{81.4 \text{ g}} \times \frac{1 \text{ mole CO}}{1 \text{ mole ZnO}} \times 28.0 \text{ g/mole} = 4.20 \text{ g CO}$$

4. (*a*) 1 mole Zn = 65.4 g

$$Y = 12.2 \text{ g ZnO} \times \frac{\text{mole}}{81.4 \text{ g}} \times \frac{1 \text{ mole Zn}}{1 \text{ mole ZnO}} \times 65.4 \text{ g/mole} = 9.81 \text{ g Zn}$$

(*b*) 1 mole CO_2 = 22.4 liters

$$Z = 12.2 \text{ g ZnO} \times \frac{\text{mole}}{81.4 \text{ g}} \times \frac{1 \text{ mole } CO_2}{1 \text{ mole ZnO}} \times 22.4 \text{ liters/mole} = 3.36 \text{ liters}$$

5. $$\underset{2 \text{ moles}}{\overset{1.00 \text{ kg}}{2NaHCO_3}} + \underset{1 \text{ mole}}{\overset{X}{H_2SO_4}} \rightarrow Na_2SO_4 + 2H_2O + \underset{2 \text{ moles}}{\overset{Y}{2CO_2(g)}}$$

1 mole $NaHCO_3$ = 84.0 g 1 mole H_2SO_4 = 98.1 g

$$X = 1.00 \text{ kg} \times \frac{10^3 \text{ g}}{\text{kg}} \times \frac{1 \text{ mole } NaHCO_3}{84.0 \text{ g}} \times \frac{1 \text{ mole } H_2SO_4}{2 \text{ moles } NaHCO_3} \times 98.1 \text{ g/mole}$$

$$= 584 \text{ g } H_2SO_4$$

6. 1 mole CO_2 = 22.4 liters

$$Y = 1.00 \text{ kg} \times \frac{10^3 \text{ g}}{\text{kg}} \times \frac{1 \text{ mole } NaHCO_3}{84.0 \text{ g}} \times \frac{2 \text{ moles } CO_2}{2 \text{ moles } NaHCO_3} \times 22.4 \text{ liters/mole}$$

$$= 267 \text{ liters } CO_2$$

Group B

7. $$\underset{1 \text{ mole}}{\overset{230 \text{ g}}{HCOOH}} \rightarrow H_2O + \underset{1 \text{ mole}}{\overset{X}{CO}}$$

1 mole HCOOH = 46.0 g 1 mole CO = 28.0 g

$$X = 230 \text{ g HCOOH} \times \frac{\text{mole}}{46.0 \text{ g}} \times \frac{1 \text{ mole CO}}{1 \text{ mole HCOOH}} \times 28.0 \text{ g/mole} = 140 \text{ g CO}$$

8. $$140 \text{ g CO} \times \frac{22.4 \text{ liters}}{28.0 \text{ g}} \times \frac{300^\circ \text{K}}{273^\circ \text{K}} \times \frac{760 \text{ mm}}{750 \text{ mm}} = 125 \text{ liters CO}$$

9. $2CO(g) + O_2 \rightarrow 2CO_2(g)$
   2 moles 2 moles

   $$125 \text{ liters CO} \times \frac{2 \text{ moles } CO_2}{2 \text{ moles CO}} = 125 \text{ liters } CO_2$$

10. $$V_{STP} = 2.50 \text{ liters} \times \frac{273^\circ K}{29\bar{0}^\circ K} \times \frac{74\bar{0} \text{ mm}}{76\bar{0} \text{ mm}} = 2.29 \text{ liters at STP}$$

    X 2.29 liters
    $CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2O + CO_2(g)$
    1 mole 2 moles 1 mole

    1 mole $CaCO_3$ = 100.1 g 1 mole CO_2 = 22.4 liters

    $$X = 2.29 \text{ liters } CO_2 \times \frac{\text{mole}}{22.4 \text{ liters}} \times \frac{1 \text{ mole } CaCO_2}{1 \text{ mole } CO_2} \times 100.1 \text{ g/mole} = 10.2 \text{ g } CaCO_3$$

11. 1 mole HCl = 36.5 g

    $$2.29 \text{ liters } CO_2 \times \frac{\text{mole}}{22.4 \text{ liters}} \times \frac{2 \text{ moles HCl}}{1 \text{ mole } CO_2} \times 36.5 \text{ g/mole} \times \frac{100\%}{38.0\%} \times \frac{1 \text{ ml}}{1.20 \text{ g}} =$$

    16.4 ml conc. HCl

12. C: $\frac{52.9 \text{ g C}}{12.0 \text{ g/mole}} = 4.41$ moles C

    O: $\frac{47.1 \text{ g O}}{16.0 \text{ g/mole}} = 2.94$ moles O

    C: O = 4.41/2.94 : 2.94/2.94 = 1.50 : 1.00 = 3.00 : 2.00

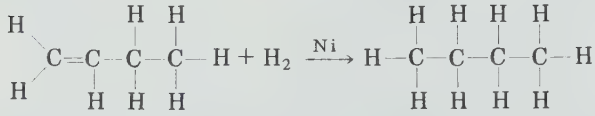
    Empirical formula C_3O_2
    $C_3O_2 + O_2 \rightarrow CO_2$ (not balanced)
    $C_3O_2 + 2O_2 \rightarrow 3CO_2$ (balanced according to conservation of atoms)
    Since the balanced equation shows 2 moles of O_2 for each mole of C_3O_2 and the volume relations are also known to be 2 volumes of O_2 for each volume of C_3O_2, C_3O_2 is proved to be the molecular formula by Avogadro's principle.

Chapter 17

Answers to Questions pp. 357–359

Group A

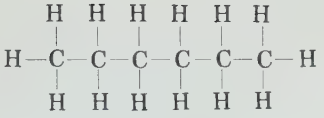
1. Carbon atoms link together with covalent bonds and the same atoms may be arranged in several different isomeric structures.
2. A pair of shared electrons.
3. The kinds of atoms, how many of each, and their arrangement in the molecule.
4. Most organic compounds do not dissolve in water. Organic compounds are decomposed by heat more easily than most inorganic compounds. Organic reactions generally proceed at much slower rates. Organic compounds exist as molecules which consist of atoms joined by covalent bonds.
5. (*a*) C_nH_{2n+2}. (*b*) C_nH_{2n}. (*c*) C_nH_{2n-2}.
6. (*a*) C_6H_{14}. (*b*) C_6H_{12}. (*c*) C_6H_{10}.
7. By using the Greek or Latin prefixes which show the number of carbon atoms in the molecule and the suffix -ane.
8. A generalized formula for an alkane.

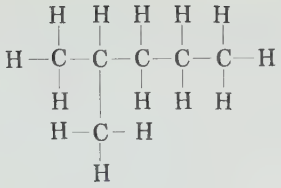
9. (*a*) Fractional distillation is a method of separating the components of a mixture on the basis of differences in their boiling points. (*b*) It is possible to separate effectively the lower members of the alkane series from petroleum and natural gas by fractional distillation. However, it is usually only practicable to separate the higher boiling substances into mixtures containing compounds having similar boiling points.
10. By heating a mixture of soda lime and sodium acetate.
11. (*a*) $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$. (*b*) $C_2H_4 + 3O_2 \rightarrow 2CO_2 + 2H_2O$. (*c*) $2C_2H_2 + 5O_2 \rightarrow 4CO_2 + 2H_2O$. (*d*) $2C_4H_6 + 11O_2 \rightarrow 8CO_2 + 6H_2O$.
12. $CH_4 + Br_2 \rightarrow CH_3Br + HBr$; $CH_3Br + Br_2 \rightarrow CH_2Br_2 + HBr$; $CH_2Br_2 + Br_2 \rightarrow CHBr_3 + HBr$; $CHBr_3 + Br_2 \rightarrow CBr_4 + HBr$.
13. 
14. (*a*) A single unit of a polymer. (*b*) A compound formed by two simpler molecules or radicals. (*c*) A compound formed by two or more simpler molecules or radicals with repeating structural units. (*d*) A union of two or more similar structures in which the atoms remain in similar relative position.
15. It is used in the synthesis of complex organic compounds and in the oxyacetylene torch.
16. They refer to the bonding in hydrocarbon molecules. Saturated molecules have only single bonds between carbon atoms, while unsaturated molecules have at least one double or triple bond.
17. Because any moisture in the air would react with the calcium carbide, and liberate ethyne.
18. They have structures which consist of two and three benzene molecules joined together, respectively.
19. Petroleum is mostly a complex mixture of hydrocarbons which varies greatly in composition depending on its source. The hydrocarbon molecules in petroleum contain from one to more than 50 carbon atoms. The types of hydrocarbons found in petroleum include alkanes, aromatic hydrocarbons, and naphthenes.
20. Petroleum is separated into suitable portions by fractional distillation.
21. (*a*) Solvent and dry cleaner. (*b*) Furnace fuel, Diesel engine fuel, cracking processes for gasoline production.
22. (*a*) Knocking occurs in an automobile engine when the mixture of gasoline vapor and air within the cylinders does not burn at a uniform rate. (*b*) By using a large proportion of molecules of branched-chain and ring hydrocarbons which burn more uniformly. Lead tetraethyl is sometimes added also.
23. The gasoline performs in a test engine like a mixture of 96% iso-octane and 4% n-heptane.
24. By adding acetic acid.
25. The monomers are joined in a zigzag chain.
26. By mixing the rubber with sulfur and various other powders.
27. Accelerators are organic catalysts which speed vulcanization. Anti-oxidants check hardness and brittleness.
28. The heating of rubber with other materials to a definite temperature for a definite time to improve its properties.
29. It is little affected by oils and greases.

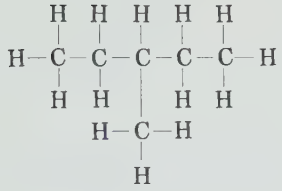
Group B

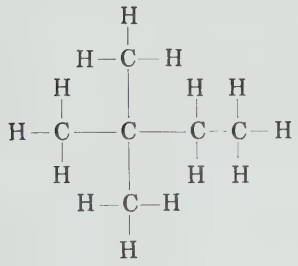
30. Ionic bonds are usually stronger than covalent bonds and more energy (higher temperature) is needed to break them.

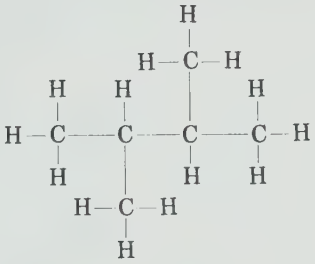
31.





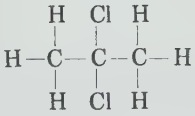






32. $2C_{10}H_{22} + 31O_2 \rightarrow 20CO_2 + 22H_2O$.

33.



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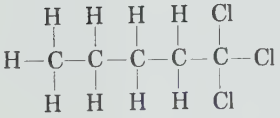
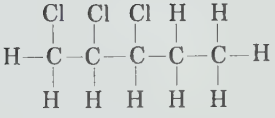
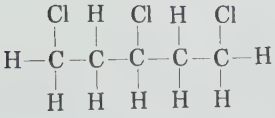


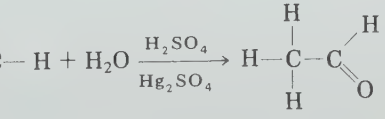




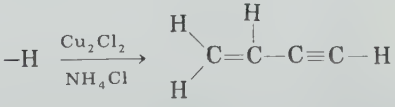
etc.

35. Almost all carbon compounds contain hydrogen. In addition hydrogen is contained in acids, hydroxides, ammonium compounds, etc.

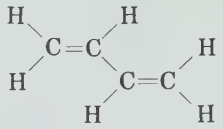
36. (*a*)



(*b*)

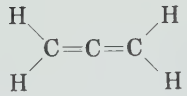


37. (*a*) Yes. (*b*) Yes. The structural formula would be

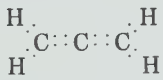


(*c*) The six hydrogen atoms and the four carbon atoms would all lie in the same plane.

38. (*a*)



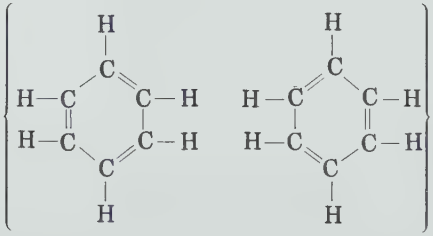
(*b*)



(*c*) Student drawing.
(*d*) They do not lie in the same plane.

39. C_6H_{2n-6}.

40. (*a*) Resonance is an attempt to make up for the deficiencies of the valence-bond structures of certain molecules written in accord with the valence-bond rules. (*b*) Benzene is represented by the resonance formula

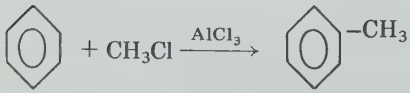


in which the two structures contribute equally. The bonds in benzene are resonance hybrid bonds. All the carbon-carbon bonds in the molecule are equivalent. (*c*) No.

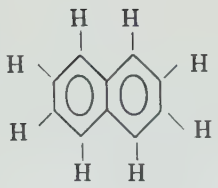
41. In ethane, the carbon-carbon bonding is sp^3-sp^3 bonding. In ethene and in benzene, it is sp^2-sp^2 bonding with *p*-*p* overlap bonding. In ethyne, it is *sp*-*sp* bonding with two *p*-*p* overlap bonds.

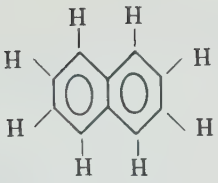
42. $C_6H_6 + Br_2 \rightarrow C_6H_5Br + HBr$; $C_6H_5Br + Br_2 \rightarrow C_6H_4Br_2 + HBr$; $C_6H_4Br_2 + Br_2 \rightarrow C_6H_3Br_3 + HBr$; $C_6H_3Br_3 + Br_2 \rightarrow C_6H_2Br_4 + HBr$; $C_6H_2Br_4 + Br_2 \rightarrow C_6HBr_5 + HBr$; $C_6HBr_5 + Br_2 \rightarrow C_6Br_6 + HBr$.

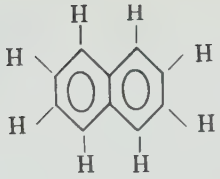
43.



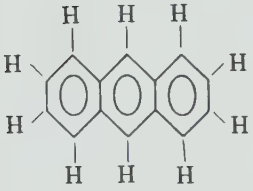
44. (*a*)

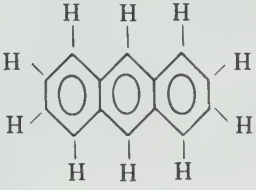


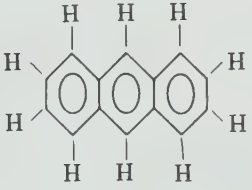


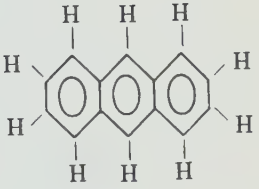


(*b*)









45. In Figure 11-6 a pure liquid is shown. In Figure 17-13 a mixture of liquids is shown. In both cases, whether the vapor bubble reaches the surface or collapses depends on the relationship between the vapor pressure of the bubble and the pressure on the liquid surface. The vapor pressure depends on the temperature.
46. Thermal cracking, increases yield; catalytic cracking, increases yield and octane rating; polymerization, increases yield; reforming, increases octane rating; alkylation, increases yield and octane rating.
47. (*a*) Butadiene and styrene. (*b*) Tire treads. (*c*) It resists wear better than the other synthetic rubbers.

Answers to Problems p. 359

Group A

1. $D = m/V$. $m = DV$. $m = 0.692$ g/ml $\times$ 15.0 gal $\times$ 4 qt/gal $\times$ 945 ml/qt $\times$ 1 kg/1000 g. $m = 39.3$ kg.
2. Molecular weight of $C_4H_{10} = 4(12.0) = 10(1.0) = 58.0$

$$\text{C: } \frac{48.0 \text{ C}}{58.0 \text{ } C_4H_{10}} \times 100\% \text{ } C_4H_{10} = 82.8\% \text{ C}$$

$$\text{H: } \frac{10.0 \text{ H}}{58.0 \text{ } C_4H_{10}} \times 100\% \text{ } C_4H_{10} = 17.2\% \text{ H}$$

3. $$\underset{1 \text{ mole}}{25.0 \text{ liters } C_3H_8} + 5O_2 \rightarrow \underset{3 \text{ moles}}{X\ 3CO_2} + 4H_2O$$

$$X = 25.0 \text{ liters } C_3H_8 \times \frac{3 \text{ moles } CO_2}{1 \text{ mole } C_3H_8} = 75.0 \text{ liters } CO_2$$

4. $\underset{1\text{ mole}}{\overset{X}{CaC_2}} + 2H_2O \rightarrow \underset{1\text{ mole}}{\overset{2.0\text{ liters}}{C_2H_2}} + Ca(OH)_2$

1 mole CaC_2 = 64.1 g 1 mole C_2H_2 = 22.4 liters

$$X = 2.0\text{ liters } C_2H_2 \times \frac{1\text{ mole}}{22.4\text{ liters}} \times \frac{1\text{ mole } CaC_2}{1\text{ mole } C_2H_2} \times \frac{64.1\text{ g}}{\text{mole}} = 5.7\text{ g } CaC_2$$

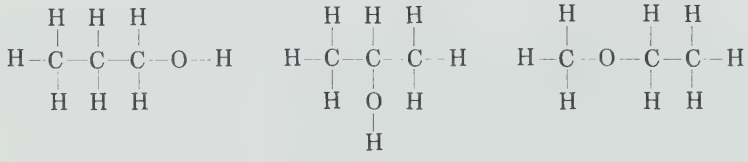
Group B

5. C: $\dfrac{60.0\text{ g C}}{12.0\text{ g/mole}} = 5.00\text{ moles C}$

O: $\dfrac{26.7\text{ g O}}{16.0\text{ g/mole}} = 1.67\text{ moles O}$

H: $\dfrac{13.3\text{ g H}}{1.0\text{ g/mole}} = 13.3\text{ moles H}$

C: O :H = 5.00/1.67:1.67/1.67:13.3/1.67 = 2.99:1.00:7.96
Empirical formula C_3OH_8
$(C_3OH_8)_x = 60$
$x = 60 \div 60.0 = 1.0$
Molecular formula = $(C_3OH_8)_1$ or C_3OH_8



6. 1 mole CH_4 + 98 kcal → 1 mole CH_3- + 1 mole H
1 mole Br_2 + 46 kcal → 2 moles Br
1 mole CH_3 + 1 mole Br → 1 mole CH_3Br + 66 kcal
1 mole H + 1 mole Br → 1 mole HBr + 87 kcal

1 mole CH_4 + 1 mole Br_2 → 1 mole CH_3Br + 1 mole HBr + 9 kcal

Exothermic energy change; 9 kcal/mole CH_3Br formed

7. $V'_{STP} = V \times p/p' \times T'/T$. $V'_{STP} = 135\text{ ml} \times \dfrac{738\text{ mm} - 20\text{ mm}}{760\text{ mm}} \times \dfrac{273^\circ\text{ K}}{295^\circ\text{ K}}$.

V'_{STP} = 118 ml.

8. (*a*)

C: $\dfrac{93.75\text{ g C}}{12.0\text{ g/mole}} = 7.81\text{ moles C}$

H: $\dfrac{6.25\text{ g H}}{1.0\text{ g/mole}} = 6.25\text{ moles H}$

C : H = 7.81/6.25 : 6.25/6.25 = 1.25 : 1.00 = 5 : 4

Empirical formula C_5H_4

(*b*)

$$\text{g-mol wt} = \frac{5.1\ C^\circ \times 6.40\ \text{g solute}}{\text{mole solute/kg } C_6H_6 \times 2.55\ C^\circ \times 10^{-1}\ \text{g } C_6H_6 \times \text{kg}/10^3\ \text{g}}$$

g-mol wt = 128 g/mole

mol wt = 128

(*c*) $(C_5H_4)_x = 128$

$x = 128 \div 64 = 2.0$

Molecular formula = $(C_5H_4)_2$ or $C_{10}H_8$

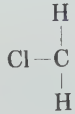
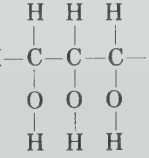
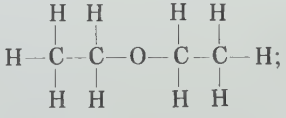
Chapter 18

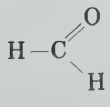
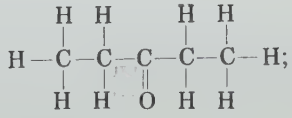
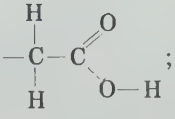
Answers to Questions pp. 373-374

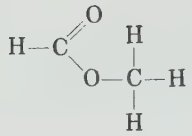
Group A

1. (*a*) RX; (*b*) ROH; (*c*) ROR′; (*d*) RCHO; (*e*) RCOR′; (*f*) RCOOH; (*g*) RCOOR′.
2. (*a*) As a solvent for grease and as a starting material for preparing other halogenated hydrocarbons. (*b*) Good ventilation is needed.
3. Alcohols do not yield hydroxide ions in water.
4. It destroys the cells of the optic nerve causing blindness and eventually death.
5. It is hygroscopic.
6. Sodium reacts with water to yield hydrogen and sodium hydroxide. Sodium reacts with methanol to yield hydrogen and sodium methoxide.
7. Six.
8. For preserving anatomical specimens.
9. (*a*) The reduction of alkaline copper(II) sulfate solution to a brick-red precipitate of copper(I) oxide. (*b*) The formyl group.
10. As a solvent in the manufacture of acetate rayon, as a cleaning agent, and in preparing synthetic organic chemicals.
11. (*a*) Vinegar made from apple cider. (*b*) By permitting apple cider to ferment to hard cider, and oxidizing the ethanol in the hard cider to acetic acid. Oxygen of air acts as the oxidizing agent. The reaction is catalyzed by enzymes from certain bacteria.
12. (*a*) $H_2C_2O_4 + H_2O \rightleftarrows H_3O^+ + HC_2O_4^-$
    $HC_2O_4^- + H_2O \rightleftarrows H_3O^+ + C_2O_4^{--}$
    (*b*) Eight moles.
13. The oxygen and one hydrogen come from the acid. The other hydrogen comes from the alcohol.
14. Dehydration of alcohols to ethers or alkynes. Esterficiation.
15. (*a*) They are esters of glycerol and long-carbon-chain acids. (*b*) A fat is a solid at room temperature while an oil is a liquid. (*c*) Oils usually contain hydrocarbon chains which are more unsaturated and have lower melting points.
16. Saponification is the alkaline hydrolysis of a fat.

Group B

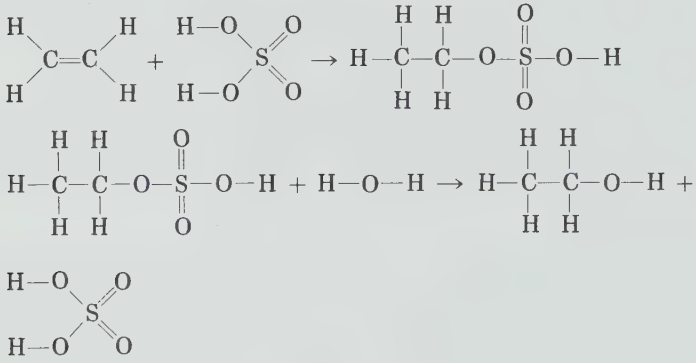
17. (*a*)  ; (*b*)  ; (*c*)  ;

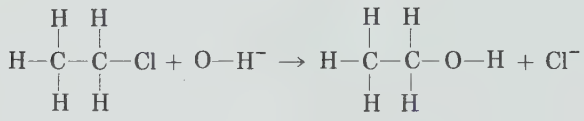
(*d*)  ; (*e*)  ; (*f*)  ;

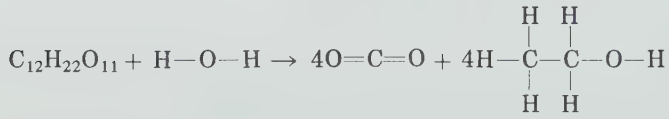
(*g*) 

18. (*a*) $C_2H_6 + Cl_2 \rightarrow C_2H_5Cl + HCl$
    (*b*) $C_2H_4 + C_2H_5Cl$
    (*c*) $C_2H_5OH + HCl \rightarrow C_2H_5Cl + H_2O$
19. (*a*) By the addition of an alkyl halide to dry ether containing magnesium. (*b*) An alcohol in which the hydroxyl group is attached to a carbon atom to which two different alkyl groups may be bonded. (*c*) An alcohol in which the hydroxyl group is attached to a carbon atom to which three different alkyl groups may be bonded.
20. $CH_4 + 4Cl_2 \rightarrow CCl_4 + 4HCl$

    $CCl_4 + 2HF \xrightarrow{\text{catalyst}} CCl_2F_2 + 2HCl$

21. (*a*) 

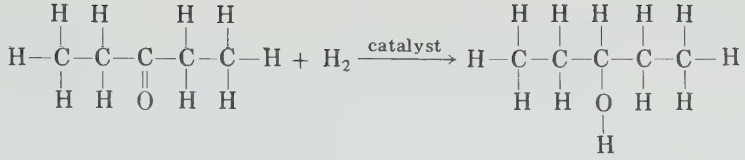
    (*b*) 

    (*c*) 

22. In order of increasing molecular weight, methanol, ethanol, ethylene glycol, increasing masses are required per unit mass of water to produce the same freezing point depression. However, in this same order, the boiling points increase. So while a smaller mass of methanol is required for a given freezing point lowering, it evaporates most readily. A larger mass of ethylene glycol is required, but it does not evaporate appreciably and will last an entire season.
23. $C_3H_7OH + HI \rightarrow C_3H_7I + H_2O$
    $3C_3H_7OH + PI_3 \rightarrow 3C_3H_7I + H_3PO_3$
24. By dehydration of propyl alcohol. By reaction of sodium propoxide with propyl bromide.
25. $C_2H_2 + H_2O \xrightarrow[Hg_2SO_4]{H_2SO_4} CH_3CHO$

    $CH_3CHO + O$ (from oxidizing agent) $\rightarrow CH_3COOH$

26. Methylethylketone.

27. 

28. The hydrogen on the carbon atom adjacent ot the carbonyl group of one aldehyde molecule bonds with the oxygen of the second aldehyde molecule. This permits the carbon of the carbonyl group of the second aldehyde molecule to bond to the carbon adjacent to the carbonyl group of the first aldehyde molecule.

29. $C_4H_9OH + CH_3COOH \xrightarrow{H_2SO_4} CH_3COOC_4H_9 + H_2O$

30. $(C_{17}H_{35}COO)_3C_3H_5 + 3NaOH \rightarrow 3C_{17}H_{35}COONa + C_3H_5(OH)_3$

Answers to Problems p. 374

Group A

1. Molecular weight of $CCl_2F_2 = 12.0 + 2(35.5) + 2(35.5) + 2(19.0) = 121.0$
2. $V_aN_a = V_bN_b$. $N_b = V_aN_a/V_b$. $N_b = (35.0 \text{ ml} \times 0.150\text{-}N)/40.0 \text{ ml}$. $N_b = 0.131\text{-}N$
3. G-mol wt $C_3H_5(OH)_3 = 92.0$

$$\frac{0.400 \text{ mole } C_3H_5(OH)_3}{\text{kg } H_2O} \times \frac{92.0 \text{ g}}{\text{mole}} \times 0.300 \text{ kg } H_2O = 11.0 \text{ g}$$

4. G-equiv wt $(COOH)_2 = \frac{\text{g-mol wt } (COOH)_2}{2} = \frac{90.0 \text{ g}}{2} = 45.0 \text{ g}$

$$\frac{0.200 \text{ g-eq } (COOH)_2}{\text{liter}} \times \frac{45.0 \text{ g}}{\text{g-eq}} \times 1.50 \text{ liters} = 13.5 \text{ g}$$

Group B

5. Formula weight of $C_{12}H_{25}OSO_2ONa = 12(12.0) + 25(1.0) + 4(16.0) + 32.1 + 23.0 = 288.1$

C: $\frac{144.0 \text{ C}}{288.1 \text{ cpd}} \times 100\% \text{ cpd} = 49.98\% \text{ C}$

H: $\frac{25.0 \text{H}}{288.1 \text{ cpd}} \times 100\% \text{ cpd} = 8.68\% \text{ H}$

O: $\frac{64.0 \text{ O}}{288.1 \text{ cpd}} \times 100\% \text{ cpd} = 22.2\% \text{ O}$

S: $\frac{32.1 \text{ S}}{288.1 \text{ cpd}} \times 100\% \text{ cpd} = 11.1\% \text{ S}$

Na: $\frac{23.0 \text{ Na}}{288.1 \text{ cpd}} \times 100\% \text{ cpd} = 7.98\% \text{ Na}$

6. (*a*) $\text{C} = \frac{62.1 \text{ g C}}{12.0 \text{ g/mole}} = 5.18 \text{ moles}$

$$H = \frac{10.3 \text{ g H}}{1.0 \text{ g/mole}} = 10.3 \text{ moles}$$

$$O = \frac{27.6 \text{ g O}}{16.0 \text{ g/mole}} = 1.72 \text{ moles}$$

C : H : O = 5.18/1.72 : 10.3/1.72 : 1.72/1.72 = 3.01 : 5.99 : 1.00
Empirical formula = C_3H_6O

(*b*) $(C_3H_6O)_x = 116$
$x = 116 \div 58 = 2$
Molecular formula = $(C_3H_6O)_2 = C_6H_{12}O_2$. The compound could be diacetone alcohol.

7. $pH = -\log [H_3O^+] = -\log (9.4 \times 10^{-4}) = -(\log 9.4 + \log 10^{-4}) = -(0.97 - 4) = 3.03$

8. G-mol wt C_2H_5OH = 46.0 g

$$\frac{0.750 \text{ mole } C_2H_5OH}{\text{liter } H_2O} \times \frac{46.0 \text{ g}}{\text{mole}} \times 500 \text{ ml} \times \frac{1 \text{ liter}}{1000 \text{ ml}} \times \frac{1 \text{ ml}}{0.789 \text{ g}} = 21.9 \text{ ml}$$

Chapter 19
Answers to Questions pp. 400-401

Group A

1. A definite amount of energy is released or absorbed when a substance reacts to form new substances even though both reactant and product are at the same temperature.
2. (*a*) Lower. (*b*) Higher.
3. The molar heat of formation of a compound is the heat evolved or absorbed when 1 mole of the compound is formed from its elements.
4. Two factors in the driving force of reactions are: the tendency toward lowest state of energy, and the tendency toward higher state of disorder.
5. In an exothermic system the heat content of the products is lower than that of the reactants, so the heat of reaction is lost to the system and the heat content changes in a negative sense.
6. Changes of state in the direction of the solid state are accompanied by decreases in entropy.
7. Rise in temperature favors an increase in entropy.
8. Activation energy is the energy required to transform reactants into the activated complex.
9. The activation energy required for the endothermic change is greater than that required for the exothermic change by the amount of the energy of reaction of the system.
10. The collision may not be energetic enough to supply the necessary activation energy, and the colliding molecules may not be oriented in a way that favors their interaction.
11. The term "activated complex" refers to the transitional structure resulting from an effective collision between reactant particles which persists while old bonds are breaking and new bonds are forming.

Group B

12. The energy change in a reaction system is related to the change that occurs in the number and strengths of bonds during the reaction.
13. 43.8 kcal – 3.0 kcal = 40.8 kcal = ΔH_a(HI)
14. A high negative heat of formation suggests that the compound is very stable. The heat content of the compound is much lower than that of its separate elements, so a large energy input is required for its decomposition.
15. If the temperature of the gases remained constant throughout the mixing, this would indicate that the heat content did not change.
16. The self-mixing of the two gases is accompanied by an increase in entropy, a driving force in natural processes.
17. If an exothermic reaction is accompanied by a decrease in entropy such that the unfavorable change in entropy, expressed as $T\Delta S$, exceeds the favorable change in energy content the reaction will not be spontaneous. This is to say, the free energy change is positive and the reaction is not spontaneous.
18. An endothermic reaction in which the entropy change is favorable and the increase in entropy, expressed as $T\Delta S$, exceeds the unfavorable energy change will proceed spontaneously since it is accompanied by a decrease in free energy.
19. (*a*) This is the pathway by which the potential energy of the reactants is raised to the minimum level needed for effective collisions. (*b*) The maximum energy region of the minimum energy pathway for reaction is reached when the activated complex is formed.
20. Since the formation of a single activated complex by the simultaneous collision of 5 molecules is highly improbable, the reaction mechanism must involve some sequence of simple steps.
21. (*a*) The rate is increased. (*b*) The rate is doubled. (*c*) The rate is cut in half. (*d*) The rate is increased. (*e*) The rate is increased.
22. $NO_2 + O \rightarrow NO + O_2$
23. (*a*) Rapid, gas chain reaction (*b*) Rapid, ionic, no bonds to break, limited only by addition rate. (*c*) Slow, limited surface area of contact.
24. Measure pressure of the reaction system. Since 2 moles of gas reactant form 1 mole of gas product, the pressure will decrease as the reaction proceeds.
25. $O_3 \rightarrow O_2 + O$
    $O + O_3 \rightarrow 2O_2$

Solutions to Problems pp. 401-402

Group A

1. $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l) + 212.80$ kcal
   $\Delta H_{f(CH_4)} = \Delta H_{f(CO_2)} + 2\Delta H_{f(H_2O)} - \Delta H_{c(CH_4)}$
   $\Delta H_{f(CH_4)} = -94.05$ kcal + (–136.64 kcal) – (–212.80 kcal) = –17.89 kcal/mole
2. $2C_2H_2(g) + 5O_2(g) \rightarrow 4CO_2(g) + 2H_2O(l) + 621.24$ kcal
   $C_2H_2(g) + \frac{5}{2}O_2(g) \rightarrow 2CO_2(g) + H_2O(l) + 310.62$ kcal
   $\Delta H_{f(C_2H_2)} = 2\Delta H_{f(CO_2)} + \Delta H_{f(H_2O)} - \Delta H_{c(C_2H_2)}$
   $\Delta H_{c(C_2H_2)} = -188.10$ kcal + (–68.32) kcal – (–310.62 kcal) = +54.20 kcal/mole
3. $2C_6H_6(l) + 15O_2(g) \rightarrow 12CO_2(g) + 6H_2O(l) + 1561.96$ kcal
   $C_6H_6(l) + \frac{15}{2}O_2(g) \rightarrow 6CO_2(g) + 3H_2O(l) + 780.98$ kcal
   $\Delta H_{f(C_6H_6)} = 6\Delta H_{f(CO_2)} + 3\Delta H_{f(H_2O)} - \Delta H_{c(C_6H_6)}$
   $\Delta H_{f(C_6H_6)} = -564.30$ kcal + (–204.96) kcal – (–780.98) kcal = +11.72 kcal/mole
4. $2H_2(g) + O_2(g) \rightarrow 2H_2O(l) + 2\Delta H_c$
   $H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l) + \Delta H_c$
   $\Delta H_{c(H_2)} = \Delta H_{f(H_2O)} - \Delta H_{f(H_2)}$
   $\Delta H_{c(H_2)} = -68.32$ kcal – 0.00 kcal = –68.32 kcal/mole

5. (*a*) $0.0375\ M - 0.0268\ M = 0.0107\ M$
0.107 mole/liter $\div$ 18.0 min $= 5.95 \times 10^{-4}$ mole/liter/min
(*b*) 5.98×10^{-4} mole/liter/min $\div$ 60 sec/min $= 9.92 \times 10^{-6}$ mole/liter/sec

Group B

6.
$$\begin{array}{llll} S(s) + O_2(g) \rightarrow SO_2(g) & \Delta H = -70.96 \text{ kcal} & (1) \\ SO_2(g) + 1/2O_2(g) \rightarrow SO_3(g) & \Delta H = -23.49 \text{ kcal} & (2) \\ SO_3(g) + H_2O(l) \rightarrow H_2SO_4(l) & \Delta H = -31.14 \text{ kcal} & (3) \\ H_2(g) + 1/2O_2(g) \rightarrow H_2O(l) & \Delta H = -68.32 \text{ kcal} & (4) \\ \hline S(s) + 2O_2(g) + H_2(g) \rightarrow H_2SO_4(l) & \Delta H = -193.91 \text{ kcal} \end{array}$$

7. (*a*) (1) $\dfrac{-57.8 \text{ kcal}}{18.0 \text{ g reactants}} \times \dfrac{10^3 \text{ g}}{\text{kg}} = \dfrac{3210 \text{ kcal}}{\text{kg reactants}}$

(2) $\dfrac{-64.2 \text{ kcal}}{20.0 \text{ g reactants}} \times \dfrac{10^3 \text{ g}}{\text{kg}} = \dfrac{3210 \text{ kcal}}{\text{kg reactants}}$

(*b*) Reaction (1) preferred since mol wt H_2O is lower

8. (*a*) $R = k[A]\,[B]^2$

(*b*) $k = \dfrac{R}{[A]\,[B]^2} = \dfrac{2.0 \times 10^{-4} \text{ mole/liter/min}}{0.20 \text{ mole/liter} \times (0.20 \text{ mole/liter})^2}$

$k = 2.5 \times 10^{-2}$ liter^2/mole^2/min

(*c*) $R = k[A]\,[B]^2$

$$R = \frac{2.5 \times 10^{-2} \text{ liter}^2}{\text{mole}^2 \text{ min}} \times \frac{0.30 \text{ mole}}{\text{liter}} \times \left(\frac{0.30 \text{ mole}}{\text{liter}}\right)^2$$

$R = 6.8 \times 10^{-4}$ mole/liter/min (or M/min)

Chapter 20
Answers to Questions pp. 429-431

Group A

1. (*1*) Water in contact with its saturated vapor in a closed vessel. (*2*) Sugar crystals in contact with a saturated solution. (*3*) Sodium chloride crystals in contact with a saturated solution. (Other examples of physical equilibrium may be given.)
2. $HC_2H_3O_2 + H_2O \rightleftarrows H_3O^+ + C_2H_3O_2^-$
$H_2CO_3 + H_2O \rightleftarrows H_3O^+ + HCO_3^-$.
$H_3PO_4 + H_2O \rightleftarrows H_3O^+ + H_2PO_4^-$.
(Other examples may be given using soluble polar compounds recognized as weak electrolytes.)
3. Equilibrium is a dynamic state; the opposite reactions do not stop but continue at the same rate so that the net change is zero.
4. The speed of a reaction is proportional to the product of the concentrations of the reacting substances, each raised to the appropriate power.
5. (*a*) The speed of the reaction increases decidedly. (*b*) Air is one-fifth oxygen; thus pure oxygen at the same pressure as the air will provide 5 times the number of collisions between the reactants.

6. (*a*) If a system at equilibrium is subjected to a stress, the equilibrium will be displaced in such direction as to relieve the stress. (*b*) Le Chatelier's principle applies to all dynamic equilibria, physical, chemical, and ionic.
7. (*a*) (*1*) Concentration. (*2*) Pressure. (*3*) Temperature. (*b*) A change in the temperature of the system causes a change in the value of the equilibrium constant for that system.
8. (*1*) If an insoluble product is formed.
   $Ag^+ + Cl^- \rightarrow AgCl$
   (*2*) If a gaseous product is formed.
   $H_3O^+ + HCO_3^- \rightarrow 2H_2O + CO_2$
   (*3*) If a product is very slightly ionized.
   $H_3O^+ + OH^- \rightarrow 2H_2O$
   (All spectator ions have been omitted from these equations.)
9. Salts that are sparingly soluble in water, generally described as insoluble.
10. The product of the molar concentrations of the ions of a sparingly soluble substance in saturated solution, each raised to the appropriate power.

Group B

11. Place the steam and iron in a closed container and heat. Neither the steam nor the hydrogen can escape. When the tube is opened, all 4 substances will be present.
12. The term dynamic means "active" as opposed to "static." Dynamic equilibria involve opposing processes which are occurring at the same time at the same speed.
13. $\dfrac{[CH_3OH]}{[CO]\,[H_2]^2} = K$
14. The reaction to the right is exothermic. A rise in temperature would act as a stress on the system and the equilibrium would shift in the direction of the endothermic reaction to relieve the stress. This would reduce the concentration of methanol. By lowering the temperature of the system the yield of methanol would be increased.
15. An increase in pressure would act as a stress and the equilibrium would shift in the direction which produced fewer gas molecules to relieve the stress. Increasing the pressure on the system would increase the yield of methanol.
16. $K = \dfrac{[C]}{[A]\,[B]} = \dfrac{1.0}{2.0 \cdot 3.0} = \dfrac{1.0}{6.0} = 0.17$
17. (*a*) $BaCO_3 + 2H_3O^+ \rightarrow Ba^{++} + 3H_2O + CO_2(g)$
    (*b*) $Pb^{++} + 2Cl^- \rightarrow PbCl_2(s)$ (low solubility in cold water)
    (*c*) NO REACTION
    (*d*) NO REACTION
    (*e*) $Ba^{++} + SO_4^{--} \rightarrow BaSO_4(s)$
    (*f*) NO REACTION
    (*g*) $AgC_2H_3O_2 + H_3O^+ + Cl^- \rightarrow HC_2H_3O_2 + AgCl(s) + H_2O$
    (*h*) $3Cu^{++} + 2PO_4^{---} \rightarrow Cu_3(PO_4)_2(s)$
    (*i*) $Ba^{++} + SO_4^{--} \rightarrow BaSO_4(s)$
    (*j*) $CuO + 2H_3O^+ \rightarrow Cu^{++} + 3H_2O$
18. Acetic acid is a weak electrolyte and is slightly ionized in water solution. Equilibrium is quickly established between acetic acid molecules and hydronium and acetate ions in the solution. $HC_2H_3O_2 + H_2O \rightleftarrows H_3O^+ + C_2H_3O_2^-$. Sodium acetate is an ionic salt and is therefore completely ionized in water solution. When added to the acetic acid solution the concentration of the acetate ion is increased, driving the equilibrium to the left. This removes hydronium ions from solution and increases the concentration of un-ionized acetic acid molecules. Thus the solution is less acidic and has a higher pH value.

19.

pH	$[H_3O^+]$ (moles/liter)	$[OH^-]$ (moles/liter)	$\frac{[H_3O^+]}{[OH^-]}$	Property
0	$10^0 = 1.0$	$10^{-14} = 0.00000000000001$	10^{-14}	↑
1	$10^{-1} = 0.1$	$10^{-13} = 0.0000000000001$	10^{-14}	Stronger
3	$10^{-3} = 0.001$	$10^{-11} = 0.00000000001$	10^{-14}	acid
5	$10^{-5} = 0.00001$	$10^{-9} = 0.000000001$	10^{-14}	|
7	$10^{-7} = 0.0000001$	$10^{-7} = 0.0000001$	10^{-14}	Neutral
9	$10^{-9} = 0.000000001$	$10^{-5} = 0.00001$	10^{-14}	|
11	$10^{-11} = 0.00000000001$	$10^{-3} = 0.001$	10^{-14}	Stronger
13	$10^{-13} = 0.0000000000001$	$10^{-1} = 0.1$	10^{-14}	base
14	$10^{-14} = 0.00000000000001$	$10^0 = 1.0$	10^{-14}	↓

20. (1) $K = \dfrac{[NH_3]^2}{[N_2][H_2]^3}$ (2) $K = \dfrac{[NH_3]^2}{[N_2][H_2]^3}$ (3) $K = \dfrac{[NH_3]^2}{[N_2][H_2]^3}$

(4) $K = \dfrac{[H_2S]^2}{[H_2]^2[S_2]}$ (5) $K = \dfrac{[H_2O][CO]}{[H_2][CO_2]}$ (6) $K = \dfrac{[H_2]^2[O_2]}{[H_2O]^2}$

(7) $K = \dfrac{[CO_2]^2}{[CO]^2[O_2]}$ (8) $K = \dfrac{[HI]^2}{[H_2][I_2]}$ (9) $K = \dfrac{[HI]^2}{[H_2][I_2]}$

(10) $K = \dfrac{[HI]^2}{[H_2][I_2]}$ (11) $K = \dfrac{[CO]^2}{[CO_2]}$ (12) $K = \dfrac{[Cu^{++}]}{[Ag^+]^2}$

(13) $K = \dfrac{[I]^2}{[I_2]}$ (14) $K = \dfrac{[O_2]^3}{[O_3]^2}$ (15) $K = \dfrac{[N]^2}{[N_2]}$

21. In the first two experiments $[H_2] = [I_2]$ because the data were acquired by starting with hydrogen iodide. The equilibrium equation shows that HI decomposition gives equal quantities of H_2 and I_2.

22. The $H_2 + I_2 \rightleftarrows HI$ equilibrium is not affected by changes in pressure since both sides of the equilibrium equation show the same number of gas molecules.

23. (*a*) $NH_3 + H_2O \rightleftarrows NH_4^+ + OH^-$

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

(*b*) The value of K_b is very small showing that there are comparatively few NH_4^+ and OH^- ions in the equilibrium.

24. $B^- + H_2O \rightleftarrows HB + OH^-$ (1)

$$K_h = \frac{[HB][OH^-]}{[B^-]} \quad (2)$$

From $2H_2O \rightleftarrows H_3O^+ + OH^-$ (3)

$K_w = [H_3O^+][OH^-]$ (4)

From $HB + H_2O \rightleftarrows H_3O^+ + B^-$ (5)

$$K_a = \frac{[H_3O^+][B^-]}{[HB]} \tag{6}$$

Solving Eq. (6) for [HB]: $[HB] = \frac{[H_3O^+][B^-]}{K_a}$ (7)

Subs. Eq. (7) in Eq. (2): $K_h = \frac{[H_3O^+][B^-][OH^-]}{K_a[B^-]} = \frac{[H_3O^+][OH^-]}{K_a}$ (8)

Subs. Eq. (4) in Eq. (8): $K_h = \frac{K_w}{K_a}$

25. The concentration of molecular acetic acid is high and the concentrations of H_3O^+ and $C_2H_3O_2^-$ ions are low. Equilibrium is established early, before the forward reaction has progressed very far. This shows that acetic acid is a weak electrolyte.

Solutions to Problems p. 431

Group A

1. $[H_3O^+] = 0.00040$ mole/liter $= 4.0 \times 10^{-4}$ mole/liter
   $\log 0.00040 = -4.60 = 6.60 - 10 = -3.40$
   or, $\log (4.0 \times 10^{-4}) = \log 4.0 + \log 10^{-4} = 0.60 - 4 = -3.40$

   $$pH = \log \frac{1}{[H_3O^+]} = -\log[H_3O^+] = -(-3.40) = 3.40$$

2. $[H_3O^+] = 0.002$ mole/liter $= 2 \times 10^{-3}$ mole/liter
   $\log 0.002 = -3.30 = 7.30 - 10 = -2.70$
   or, $\log (2 \times 10^{-3}) = \log 2 + \log 10^{-3} = 0.30 - 3 = -2.70$

   $$pH = \log \frac{1}{[H_3O^+]} = \log[H_3O^+] = -(-2.70) = 2.70$$

3. $[OH^-] = 0.02$ mole/liter $= 2 \times 10^{-2}$ mole/liter
   $\log 0.02 = -2.30 = 8.30 - 10 = -1.70$
   or, $\log(2 \times 10^{-2}) = \log 2 + \log 10^{-2} = 0.30 - 2 = -1.70$
   $pH = 14 - 1.70 = 12.30$
   (Alternate answer to No. 3)
   $[H_3O^+][OH^-] = 1 \times 10^{-14}$ mole/liter

   $$[H_3O^+] = \frac{1 \times 10^{-14} \text{ mole/liter}}{[OH^-]} = \frac{1 \times 10^{-14} \text{ mole/liter}}{2 \times 10^{-2} \text{ mole/liter}}$$

   $[H_3O^+] = 5 \times 10^{-13}$ mole/liter
   $\log (5 \times 10^{-13}) = \log 5 + \log 10^{-13} = 0.70 - 13 = 12.30$

   $$pH = \frac{1}{[H_3O^+]} = -\log[H_3O^+] = -(-12.30) = 12.30$$

4. (*a*) Gram-formula weight of NaOH = 40 g
   0.50-*M* solution = 0.50 mole solute/liter solution

   $$\text{Mass NaOH} = \frac{40 \text{ g NaOH}}{\text{mole}} \times \frac{0.50 \text{ mole}}{\text{liter}} \times 0.25 \text{ liter} = 5.0 \text{ g NaOH}$$

   5.0 g NaOH dissolved in water in volumetric flask and diluted with water to 250 ml volume.
   (*b*) Normality = 0.50 *N*

5. Gram-formula weight of $CuSO_4 \cdot 5H_2O = 249.7$ g
2.00-*M* solution = 2.00 moles solute/liter solution

$$\text{mass } CuSO_4 \cdot 5H_2O = \frac{249.7 \text{ g}}{\text{mole}} \times \frac{2.00 \text{ moles}}{\text{liter}} \times 0.750 \text{ liter} = 374 \text{ g}$$

Group B

6. $[H_3O^+] = [C_2H_3O_2^-] = 10^{-3.3799}$
$\text{antilog}(-3.3799) = \text{antilog}(0.6201 - 4) = 4.17 \times 10^{-4}$
$[H_3O^+] = [C_2H_3O_2^-] = 4.17 \times 10^{-4}$
$[HC_2H_3O_2] = 0.01000 - 0.000417 = 0.00958$

$$K_a = \frac{[H_3O^+][C_2H_3O_2^-]}{[HC_2H_3O_2]} = \frac{(4.17 \times 10^{-4})^2}{9.58 \times 10^{-3}}$$

$$K_a = \frac{17.4 \times 10^{-8}}{9.58 \times 10^{-3}} = 1.82 \times 10^{-5}$$

7. (*Determination of* $[OH^-]$)
$NH_3 + H_2O \rightleftarrows NH_4^+ + OH^-$

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

At equilibrium,
$[NH_4^+] = [OH^-]$
$[NH_3] \simeq 0.50\ M$ (*See problem Note*)
$K_b = 1.8 \times 10^{-5}$

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]} = \frac{[OH^-]^2}{0.50} = 1.8 \times 10^{-5}$$

$$[OH^-]^2 = 0.90 \times 10^{-5}$$

$$[OH^-] = \sqrt{9.0 \times 10^{-6}} = 3.0 \times 10^{-3}$$

(*Determination of pH*)
$K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$

$$[H_3O^+] = \frac{K_w}{[OH^-]} = \frac{1.0 \times 10^{-14}}{3.0 \times 10^{-3}}$$

$[H_3O^+] = 3.3 \times 10^{-12}$
$\text{pH} = -\log[H_3O^+] = -\log(3.3 \times 10^{-12}) = -(0.52 - 12) = 11.5$

8. *At equilibrium*
$AgBr \rightleftarrows Ag^+ + Br^-$

$$1.3 \times 10^{-4} \text{ g AgBr/liter} = \frac{1.3 \times 10^{-4} \text{ g AgBr/liter}}{188 \text{ g/mole}} = 6.9 \times 10^{-7}\ M \text{ AgBr}$$

$6.9 \times 10^{-7}\ M\ AgBr \rightarrow 6.9 \times 10^{-7}\ M\ Ag^+ + 6.9 \times 10^{-7}\ M\ Br^-$
$K_{sp} = [Ag^+][Br^-] = (6.9 \times 10^{-7})(6.9 \times 10^{-7}) = 4.8 \times 10^{-13}$

9. *At equilibrium*
$AgC_2H_3O_2 \rightleftarrows Ag^+ + C_2H_3O_2^-$
1 mole $AgC_2H_3O_2 \rightarrow 1$ mole $Ag^+ + 1$ mole $C_2H_3O_2^-$

$K_{sp} = [Ag^+][C_2H_3O_2^-] = 2.5 \times 10^{-3}$
$[Ag^+] = [C_2H_3O_2^-] = [AgC_2H_3O_2]$ dissolved
$[AgC_2H_3O_2]^2 = 2.5 \times 10^{-3}$
$[AgC_2H_3O_2] = \quad 25 \times 10^{-4} = 5.0 \times 10^{-2}\ M$

$$\frac{5.0 \times 10^{-2} \text{ mole } AgC_2H_3O_2}{\text{liter}} \times \frac{167 \text{ g}}{\text{mole}} \times 10.0 \text{ liters} = 84 \text{ g } AgC_2H_3O_2$$

10. (*a*) $PbSO_4$ might precipitate.
(*b*) $PbSO_4 \rightleftharpoons Pb^{++} + SO_4^{--}$
$K_{sp} = [Pb^{++}][SO_4^{--}] = 1.6 \times 10^{-8}$
If the $Pb(NO_3)_2$ dissolves completely,
$[Pb^{++}] = 0.0015$ mole/liter
$[SO_4^{--}] = 0.0015$ mole/liter
Ion product $[Pb^{++}][SO_4^{--}] = (1.5 \times 10^{-3})(1.5 \times 10^{-3}) = 2.2 \times 10^{-6}$
Since $2.2 \times 10^{-6} > 1.9 \times 10^{-8}$, $PbSO_4$ precipitates.

Chapter 21

Answers to Questions pp. 455–456

Group A

1. Oxidation involves the removal of electrons from particles. Reduction involves the acquisition of electrons by particles.
2. The removal and acquisition of electrons are interdependent processes. In a chemical action in which a substance gains electrons, another substance must lose a like number.
3. The oxidation state changes by algebraic addition of a negative quantity.
4. Oxidation-reduction reactions: (*a*), (*b*), (*c*), (*f*), (*g*), (*h*), and (*j*).
5.

Equation	(*a*)	(*b*)	(*c*)	(*d*)
(*a*)	sodium	chlorine	chlorine	sodium
(*b*)	carbon	oxygen	oxygen	carbon
(*c*)	oxygen	hydrogen	hydrogen (indirectly)	oxygen (indirectly)
(*f*)	oxygen	chlorine	chlorine	oxygen
(*g*)	hydrogen	chlorine	chlorine	hydrogen
(*h*)	hydrogen	oxygen	oxygen	hydrogen
(*j*)	zinc	copper(II) ion	copper(II) ion	zinc

6. (*a*) $\overset{+2}{Ca} = +2$, $\overset{+6-2}{Cl} = +4$, $\overset{+1-1}{O} = -2$. (*b*) $\overset{+1}{Na} = +1$, $\overset{+1}{H} = +1$, $\overset{+6-2}{P} = +5$, $\overset{+1+4-2}{O} = -2$. (*c*) $\overset{+1}{K} = +1$, $\overset{+1+6-2}{Si} = +4$, $\overset{+1+7-2}{O} = -2$. (*d*) $\overset{+1}{H} = +1$, $\overset{+3}{P} = +3$, $\overset{-2}{O} = -2$. (*e*) $\overset{+3}{Fe} = +3$, $\overset{-2}{O} = -2$, $\overset{+1}{H} = +1$.
7. (*a*) $PbSO_4$; (*b*) H_2O_2; (*c*) $K_2Cr_2O_7$; (*d*) H_2SO_3; (*e*) $HClO_4$.
8. Anode—lead(IV) oxide; cathode—lead; electrolyte—moderately dilute sulfuric acid solution.
9. (*a*) Electrode potential is the potential difference between an electrode and its solution in a half-reaction.
(*b*) Half-reaction is the reaction taking place at the electrode in a half-cell.
(*c*) Half-cell is the portion of a voltaic cell consisting of an electrode in contact with a solution of its ions.
10. By assigning an arbitrary potential of 0.00 volt to the standard hydrogen electrode, the potential difference across a cell using hydrogen as the reference electrode can be attributed to the other half-reaction.

Group B

11. Step 1: Write skeleton equation.
Step 2: Assign oxidation numbers.

Step 3: Write electronic equations for the oxidation and the reduction.
Step 4: Interbalance the electronic equations.
Step 5: Transfer coefficients to skeleton equation.
Step 6: Adjust coefficients of the remainder of the equation.

12. (*a*)
$$\overset{0}{Zn} + \overset{+1}{H}Cl \rightarrow \overset{+2}{Zn}Cl_2 + \overset{0}{H_2}$$
$$Zn \rightarrow Zn^{++} + 2e^-$$
$$2H^+ + 2e^- \rightarrow H_2$$

(*b*)
$$\overset{0}{Fe} + \overset{+2}{Cu}SO_4 \rightarrow \overset{+2}{Fe}SO_4 + \overset{0}{Cu}$$
$$Fe \rightarrow Fe^{++} + 2e^-$$
$$Cu^{++} + 2e^- \rightarrow Cu$$

(*c*)
$$\overset{0}{Cu} + \overset{+6}{H_2}SO_4 \rightarrow \overset{+2}{Cu}SO_4 + \overset{+4}{SO_2} + H_2O$$
$$Cu \rightarrow Cu^{++} + 2e^-$$
$$\overset{+6}{S} + 2e^- \rightarrow \overset{+4}{S}$$

(*d*)
$$\overset{-1}{H}Cl + \overset{+7}{K}MnO_4 \rightarrow \overset{+2}{Mn}Cl_2 + KCl + \overset{0}{Cl_2} + H_2O$$
$$\overset{+7}{} 10Cl^- \rightarrow 5Cl_2 + 10e^-$$
$$2Mn + 10e^- \rightarrow 2Mn^{++}$$

(*e*)
$$\overset{0}{Br_2} + H_2O \rightarrow \overset{-1}{H}Br + \overset{+1}{H}BrO$$
$$Br \rightarrow \overset{+1}{Br} + e^-$$
$$Br + e^- \rightarrow Br^-$$

13.
$$\overset{0}{Cu} + \overset{+5}{H}NO_3\ (conc.) \rightarrow \overset{+2}{Cu}(NO_3)_2 + H_2O + \overset{+4}{NO_2}$$
$$Cu \rightarrow Cu^{++} + 2e^-$$
$$\overset{+5}{2N} + 2e^- \rightarrow \overset{+4}{2N}$$
$$Cu + 4HNO_3 \rightarrow Cu(NO_3)_2 + 2H_2O + 2NO_2$$

14.
$$\overset{0}{Cu} + \overset{+5}{H}NO_3(dilute) \rightarrow \overset{+2}{Cu}(NO_3)_2 + H_2O + 2NO$$
$$3Cu \rightarrow 3Cu^{++} + 6e^-$$
$$2N + 6e^- \rightarrow \overset{+2}{2N}$$
$$3Cu + 8HNO_3 \rightarrow 3Cu(NO_3)_2 + 4H_2O + 2NO$$

15. The metals at the top of the table are strong reducing agents. They readily give up electrons to oxidized substances appearing below them. This property diminishes to the extent that metals near the bottom give up electrons only in the presence of strong oxidizing agents.

16. $\overset{+6}{K_2}\overset{+6}{Cr_2}O_7 + HCl \rightarrow KCl + \overset{+3}{Cr}Cl_3 + H_2O + \overset{0}{Cl_2}$
$Cl^{-1} \rightarrow Cl + e^-$ (multiply by 6) $6Cl^- \rightarrow 3Cl_2 + 6e^-$
$Cr^{+6} + 3e^- \rightarrow Cl^{+3}$ (multiply by 2) $2Cr^{+6} + 6e^- \rightarrow 2Cr^{+3}$
$K_2Cr_2O_7 + 14HCl \rightarrow 2KCl + 2CrCl_3 + 7H_2O + 3Cl_2$
17. $Cu^0 \rightarrow Cu^{++} + 2e^-$ $E = -0.34$ v
$Ag^+ + e^-$ Ag^0 $E = +0.80$ v
18. (*a*) -0.34 v + 0.80 v = 0.46 v
Cu electrode is negative with respect to the Ag electrode.
(*b*) $Cu + 2Ag^+ \rightarrow Cu^{++} + 2Ag$

Chapter 22
Answers to Questions pp. 465–466

Group A

1. The first element, sodium, one electron in the outer shell, is a highly reactive metal. The next two elements, magnesium and aluminum, two and three outer-shell electrons respectively, are metals of decreasing activity. The fourth element, silcion, four outer-shell electrons, is a metalloid. The fifth, sixth, and seventh elements phosphorus, sulfur, and chlorine, five, six, and seven outer-shell electrons respectively, are nonmetals of increasing activity. The eighth element, argon, eight outer-shell electrons, is an inert noble gas.
2. (*a*) Good conduction of heat and electricity, metallic luster, ductility, malleability. (*b*) Poor conduction of heat, no conduction of electricity, no luster, brittleness.
3. (*a*) Attraction between positive ions and negative electron gas. (*b*) Strong covalent bonds. (*c*) Covalent bond. (*d*) Dispersion interaction forces. (*e*) Dispersion interaction forces.
4. Aluminum has a metallic structure while silicon has a covalent network structure. The bonding forces are much stronger between the atoms of solid silicon than between the ions of solid aluminum within their electron gas.
5. The interatomic forces of elements which are malleable are stronger than those of elements which are brittle.
6. Because the sodium ion has no third-energy-level electron.
7. Because it has one excess electron and the chlorine nucleus does not hold the outer shell octet as close as it holds the seven outer-shell electrons in the neutral atom.
8. Na_2O reacts with water to produce hydrogen and sodium hydroxide solution. SiH_4 does not react with water. Cl_2O reacts with water to form a weak hypochlorous acid solution.
9. NaH reacts with water to produce hydrogen and sodium hydroxide solution. SiH_4 reacts with water to form silicon dioxide and hydrogen. HCl reacts with water to form hydronium ions and chloride ions.
10. The silicon-silicon bonds and silicon-hydrogen bonds of the silanes are weaker than the carbon-carbon bonds and carbon-hydrogen bonds of the alkanes.

Group B

11. The silvery luster of metals can be attributed to the free electrons in the crystal surfaces. When light strikes these surfaces, the valence electrons absorb energy and are set into vibratory motions that re-radiate the energy in all directions as light.
12. Because the heavier aluminum atoms are smaller and pack more closely together.
13. The binding force between atoms of liquid argon is very much smaller than the binding force between the atoms of liquid aluminum, as shown by the very much higher boiling temperature of aluminum as compared to argon.
14. There are no free valence electrons in diamond, while there are some free valence electrons in silicon.
15. By using three single covalent bonds, it is possible for four phosphorus atoms to bond to each other in a pyramidal-type structure. Sulfur atoms bond together in eight-membered puckered rings, with each atom using only two covalent bonds. Chlorine atoms can only bond together in pairs, since each atom has only one covalent bond.
16. The first electron removed from an aluminum atom structure is a $3p$ electron, while the first electron removed from a magnesium atom structure is a $3s$ electron. Since the $3p$ electron is in a higher energy sublevel, it requires less energy to remove it than it does to remove a $3s$ electron. There appears to be structural stability to half-filled sublevels. The first electron to be removed from a sulfur atom structure is the fourth $3p$ electron. The first electron to be removed from a phosphorus atom structure is the third $3p$ electron. Since a half-filled $3p$ sublevel has greater stability, it requires less energy to remove the fourth electron from it than it does to remove the third.

17. (*a*) Na· Mg: Al: ·Si: ·P: ·S: :Cl: :Ar:

(*b*) Sodium attains its oxidation state of +1 by losing its single outer-shell electron. Magnesium and aluminum attain their oxidation states of +2 and +3 respectively in the same manner. Silicon attains its oxidation state of +4 by sharing its four outer-shell electrons. Phosphorus attains its +3 oxidation state by sharing its 3*p* electrons. It attains its +5 state by sharing all its third energy level electrons. Sulfur attains its −2 oxidation state by acquiring two additional electrons to complete the octet. It attains its +4 state by sharing its 3*p* electrons and attains its +6 state by sharing all its third energy level electrons. Chlorine attains its −1 state by acquiring one electron to complete the octet. It attains its +5 state by sharing its 3*p* electrons and its +7 state by sharing all its third energy level electrons. Argon does not gain, lose, or share electrons. Its oxidation state is 0.

18.

	Formula	Electronegativity Difference	Percentage Ionic Character	Type of Bonding
(*a*)	Na_2O	3.5 − 0.9 = 2.6	82%	Ionic
	MgO	3.5 − 1.2 = 2.3	73%	Ionic
	Al_2O_3	3.5 − 1.5 = 2.0	68%	Ionic
	SiO_2	3.5 − 1.8 = 1.7	51%	Polar covalent
	P_4O_{10}	3.5 − 2.1 = 1.4	39%	Polar covalent
	SO_2	3.5 − 2.5 = 1.0	22%	Polar covalent
	Cl_2O	3.5 − 3.0 = 0.5	7%	Slightly polar covalent
(*b*)	NaH	2.1 − 0.9 = 1.2	30%	Polar covalent predicted Actually ionic
	$(MgH_2)_x$	2.1 − 1.2 = 0.9	19%	Polar covalent
	$(AlH_3)_x$	2.1 − 1.5 = 0.6	9%	Polar covalent
	SiH_4	2.1 − 1.8 = 0.3	3%	Slightly polar covalent
	PH_3	2.1 − 2.1 = 0	0%	Covalent
	H_2S	2.5 − 2.1 = 0.4	4%	Slightly polar covalent
	HCl	3.0 − 2.1 = 0.9	19%	Polar covalent

19. Elemental silicon has each silicon atom tetrahedrally bonded to four other silicon atoms. Silicon dioxide has a similar structure except that an oxygen atom is inserted into each silicon-silicon bond. Hence the separation of silicon atoms from each other is greater in silicon dioxide than in elementary silicon.
20. (*a*) Same shape: PH_3 is slightly larger; bonding in PH_3 covalent, but not polar as in NH_3. (*b*) H_2S has same shape but smaller bond angle; H_2S is slightly larger, bonding in both is polar covalent, but polarity is less in H_2S.
21. $AlH_3 + 3H_2O \rightarrow Al(OH)_3 + 3H_2$

Answers to Problems pp. 466–467

Group A

1. $D = m/V$. $m = D \times V$. $m = 2.70\ g/cm^3 \times 5.0\ cm \times 2.0\ cm \times 1.5\ cm$. $m = 40\ g$.

2. $$0.472\ g\ Mg \times \frac{mole}{24.3\ g} \times \frac{6.02 \times 10^{23}\ atoms}{mole} = 1.17 \times 10^{22}\ atoms.$$

3. Formula weight of $P_4O_{10} = 4(31.0) + 10(16.0) = 284.0$

$$P: \quad \frac{(124.0)\ P}{(284.0)\ P_4O_{10}} \times 100\%\ P_4O_{10} = 43.66\%\ P$$

4. $$\underset{1\ volume}{160\ ml\ SiH_4} + 2H_2O \rightarrow SiO_2 + \underset{4\ volumes}{X\ 4H_2}$$

$$X = 160\ ml\ SiH_4 \times \frac{4\ vols\ H_2}{1\ vol\ SiH_4} = 640\ ml\ H_2$$

5. $\underset{1\text{ mole}}{Ca_3P_2} + 6H_2O \rightarrow 3Ca(OH)_2 + \underset{2\text{ moles}}{2PH_3}$ (X above Ca_3P_2; 5.0 g above PH_3)

1 mole Ca_3P_2 = 182.3 g 1 mole PH_3 = 34.0 g

$$X = 5.0\text{ g } PH_3 \times \frac{\text{mole}}{34.0\text{ g}} \times \frac{1\text{ mole } Ca_3P_2}{2\text{ moles } PH_3} \times \frac{182.3\text{ g}}{\text{mole}} = 13\text{ g } Ca_3P_2$$

6. $\underset{1\text{ mole}}{SiO_2} + 4HF \rightarrow 2H_2O + \underset{1\text{ mole}}{SiF_4}$ (12 g above SiO_2; X above SiF_4)

1 mole SiO_2 = 60.1 g 1 mole SiF_4 = 22.4 liters

$$X = 12\text{ g } SiO_2 \times \frac{\text{mole}}{60.1\text{ g}} \times \frac{1\text{ mole } SiF_4}{1\text{ mole } SiO_2} \times \frac{22.4\text{ liters}}{\text{mole}} = 4.5\text{ liters } SiF_4$$

Group B

7. (*a*) $$\text{g-mol wt} = \frac{5.1\ C^\circ \times 0.14\text{ g solute}}{\text{mole solute/kg}C_6H_6 \times 0.58\ C^\circ \times 10.0\text{ g } C_6H_6 \times \text{kg}/10^3\text{ g}}$$

g-mol wt = 120 g
mol wt = 120

(*b*) $(P)_x = 120$
$x = 120 \div 31 = 4$
molecular formula = P_4

8. $[OH^-] = 3.1 \times 10^{-5}$, so
$[H^+] = 3.2 \times 10^{-10}$
$pH = -\log [H^+] = -\log [3.2 \times 10^{-10}] = -\log 3.2 + 10 = -0.51 + 10 = 9.49$

Chapter 23

Answers to Questions pp. 484–485

Group A

1. The atoms in this family have electron configurations like the preceding noble gases with the addition of one electron in the next higher shell. The ions have the electron configurations of the preceding noble gases.
2. All three are prepared by the electrolysis of their fused chlorides.
3. In making "Ethyl" fluid for antiknock gasolines, in making dyes, in sodium vapor lamps, as a reducing agent, etc.
4. Caustic substances are capable of changing some types of animal and vegetable matter into soluble materials by chemical action. Corrosive substances eat away by chemical action, such as the action of an acid on a metal.
5. (*a*) Common salt, limestone, and coal. (*b*) Coal, tar, sodium carbonate, sodium hydrogen carbonate, and calcium chloride.
6. (*a*) Sodium hydroxide. (*b*) Sodium carbonate. (*c*) Sodium hydrogen carbonate.
7. Molasses contains acids which liberate carbon dioxide from baking soda.
8. Large deposits of potassium chloride, crystallized with magnesium and calcium compounds, are found in Texas and New Mexico. Some potassium compounds are extracted from Searles Lake in California.
9. (*a*) Under kerosene. (*b*) Because they react with oxygen and water vapor in the air.
10. $2K(s) + 2H_2O \rightarrow 2K^+(aq) + 2OH^-(aq) + H_2(g)$
11. A clean platinum wire is dipped into a solution of either a lithium, sodium, or potassium compound and then placed in an almost colorless oxidizing flame. The lithium atoms cause a carmine flame; the sodium, a yellow flame; and the potassium, a violet flame.
12. Both metals are sensitive over the full range of the visible spectrum.
13. $CO_2(g) + NH_3(g) + H_2O \rightarrow NH_4^+(aq) + HCO_3^-(aq)$
$Na^+(aq) + HCO_3^-(aq) \rightarrow NaHCO_3(s)$
$2NaHCO_3(s) \rightarrow Na_2CO_3(s) + H_2O(g) + CO_2(g)$

14. $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$
$CaO + H_2O \rightarrow Ca(OH)_2$
$NH_4^+(aq) + OH^-(aq) \rightarrow NH_3(g) + H_2O$
15. They are used to analyze alloys during their production and to determine the composition of the stars.

Group B

16 Because the atoms in the solid metals are rather widely spaced, and are held together by only weak attractive forces.
17. Sodium chloride is essential for proper functioning of the body, but it is lost through perspiration and must be replaced.
18. It is the cheapest most readily available source of sodium.
19. (*a*) Sodium compounds are cheaper per gram, and one gram of a sodium compound goes further. (*b*) In black gunpowder and in fertilizers.
20. Calcium chloride.
21. The salt contains magnesium chloride, which is very deliquescent.
22. The increase in volume between sodium and potassium atoms is proportionally greater than the increase in mass between the two.
23. $CO_2 + 2OH^- \rightarrow CO_3^{--} + H_2O$
24. Because the sodium hydrogen carbonate is practically insoluble under the conditions of the reaction.
25. A water solution of sodium carbonate has basic properties because of hydrolysis of the CO_3^{--} ion.
26. Lithium hydroxide. You obtain more hydroxide ions per pound.
27. Common-ion effect: the presence of $NaC_2H_3O_2$ greatly increases the $[C_2H_3O_2^-]$ and shifts the $HC_2H_3O_2 + H_2O \rightarrow H_3O^+ + C_2H_3O_2^-$ equilibrium to the left. The $[H_3O^+]$ decreases and the pH rises.

Solutions to Problems p. 485

Group A.

1. (*a*) $\underset{2\text{ moles}}{2NaOH} + \underset{1\text{ mole}}{H_2SO_4} \rightarrow Na_2SO_4 + 2H_2O$ (10.0 g over NaOH, X over H_2SO_4)

$1\text{ mole NaOH} = 40.0\text{ g} \qquad 1\text{ mole } H_2SO_4 = 98.1\text{ g}$

$$X = \frac{10.0\text{ g NaOH}}{40.0\text{ g/mole}} \times \frac{1\text{ mole } H_2SO}{2\text{ moles NaOH}} \times \frac{98.1\text{ g}}{\text{mole}} = 12.3\text{ g } H_2SO_4$$

(*b*) $\underset{2\text{ moles}}{2KOH} + \underset{1\text{ mole}}{H_2SO_4} \rightarrow K_2SO_4 + 2H_2O$ (10.0 g over KOH, X over H_2SO_4)

$1\text{ mole KOH} = 56.1\text{ g} \qquad 1\text{ mole } H_2SO_4 = 98.1\text{ g}$

$$X = \frac{10.0\text{ g KOH}}{56.1\text{ g/mole}} \times \frac{1\text{ mole } H_2SO_4}{2\text{ moles KOH}} \times \frac{98.1\text{ g}}{\text{mole}} = 8.74\text{ g } H_2SO_4$$

2. $\underset{1\text{ mole}}{NaNO_3} + \underset{1\text{ mole}}{KCl} \rightarrow \underset{1\text{ mole}}{KNO_3} + NaCl$ (1.00 kg over $NaNO_3$, 1.00 kg over KCl, X over KNO_3)

$1\text{ mole } NaNO_3 = 85.0\text{ g} \qquad 1\text{ mole KCl} = 74.6\text{ g} \qquad 1\text{ mole } KNO_3 = 101.1\text{ g}$
KCl is in excess. Therefore the mass of KNO_3 produced will depend on the mass of $NaNO_3$ used.

$$X = \frac{1.00\text{ kg } NaNO_3}{85.0\text{ g/mole}} \times \frac{1\text{ mole } KNO_3}{\text{mole } NaNO_3} \times \frac{101.1\text{ g}}{\text{mole}} = 1.19\text{ kg } KNO_3$$

3. $286/106 \times 12.5¢/\text{kg} = 33.7¢/\text{kg}$.

4. (*a*) $X = 50.0 \text{ g } Na_2CO_3 \times \frac{\text{mole}}{106 \text{ g}} \times \frac{1 \text{ mole } CO_2}{\text{mole } Na_2CO_3} \times \frac{22.4 \text{ liters}}{\text{mole}} = 10.6 \text{ liters } CO_2$

(*b*) $X = 50.0 \text{ g } NaHCO_3 \times \frac{\text{mole}}{84.0 \text{ g}} \times \frac{1 \text{ mole } CO_2}{\text{mole } NaHCO_3} \times \frac{22.4 \text{ liters}}{\text{mole}} = 13.3 \text{ liters } CO_2$

(*c*) $X = 50.0 \text{ g } K_2CO_3 \times \frac{\text{mole}}{138 \text{ g}} \times \frac{1 \text{ mole } CO_2}{\text{mole } K_2CO_3} \times \frac{22.4 \text{ liters}}{\text{mole}} = 8.12 \text{ liters } CO_2$

(*d*) $X = 50.0 \text{ g } KHCO_3 \times \frac{\text{mole}}{100 \text{ g}} \times \frac{1 \text{ mole } CO_2}{\text{mole } KHCO_3} \times \frac{22.4 \text{ liters}}{\text{mole}} = 11.2 \text{ liters } CO_2$

Group B

5. $X = \frac{1.00 \times 10^3 \text{ kg } Na_2CO_3 \times 117 \text{ g NaCl}}{106 \text{ g } Na_2CO_3} = 1.10 \times \text{kg NaCl}$

6. $X = \frac{1.00 \times 10^3 \text{ kg } Na_2CO_3 \times 22.4 \text{ liters } CO_2}{106 \text{ g } Na_2CO_3} \times \frac{\text{m}^3}{10^3 \text{ liters}} \times \frac{10^3 \text{ g}}{\text{kg}} = 211 \text{ m}^3 \text{ } CO_2$

7. $V' = 211 \text{ m}^3 \text{ } CO_2 \times \frac{760 \text{ mm}}{745 \text{ mm}} \times \frac{423° \text{ K}}{273° \text{ K}} = 334 \text{ m}^3 \text{ } CO_2$

8. $\underset{\substack{\text{1 mole} \\ 100 \text{ g/mole}}}{2.50 \times 10^3 \text{ kg } CaCO_3(s)} \rightarrow \underset{\substack{\text{1 mole} \\ 56.1 \text{ g/mole}}}{X \text{ } CaO(s)} + \underset{\substack{\text{1 mole} \\ 22.4 \text{ liters/mole}}}{Y \text{ } CO_2(g)}$

$2.72 \times 10^3 \text{ kg} \times 0.920 = 2.50 \times 10^3 \text{ kg } CaCO_3$

$X = 2.50 \times 10^3 \text{ kg } CaCO_3 \times \frac{\text{mole}}{100 \text{ g}} \times \frac{1 \text{ mole CaO}}{\text{mole } CaCO_3} \times \frac{56.1 \text{ g}}{\text{mole}}$

$= 1.40 \times 10^3 \text{ kg CaO}$

9. $Y = 2.50 \times 10^3 \text{ kg } CaCO_3 \times \frac{\text{mole}}{100 \text{ g}} \times \frac{1 \text{ mole CaO}}{\text{mole } CaCO_3} \times \frac{22.4 \text{ liters}}{\text{mole}} \times \frac{10^3 \text{ g}}{\text{kg}} =$

$= 5.60 \times 10^5 \text{ liters } CO_2 \text{ (at STP)}$

$V' = \frac{5.60 \times 10^5 \text{ liter } CO_2 \times 760 \text{ mm} \times 298° \text{ K}}{855 \text{ mm} \times 273° \text{ K}} = 5.44 \times 10^5 \text{ liters } CO_2$

10. $\frac{1.1 \text{ g NaOH}}{500 \text{ ml sol'n}} \times \frac{\text{mole}}{40.0 \text{ g}} \times \frac{10^3 \text{ ml}}{\text{liter}} = 0.055 \frac{\text{mole NaOH}}{\text{liter sol'n}}$

$[H_3O^+] = \frac{1.0 \times 10^{-14} \text{ mole}^2/\text{liter}^2}{5.5 \times 10^{-2} \text{ mole/liter}} = 1.8 \times 10^{-13} \text{ mole/liter}$

$pH = -\log[H_3O^+] = -\log(1.8 \times 10^{-13})$
$pH = -\log 1.8 + \log 10^{-13}) = -(0.26 - 13) = 12.74$

Chapter 24
Answers to Questions pp. 499-500

Group A

1. A beryllium-copper alloy.
2. For asbestos shingles, for covering steam pipes and furnaces, for automobile brake linings, for fireproof curtains and clothing, and for insulating electric appliances.
3. (*a*) A double carbonate for magnesium and calcium. (*b*) As a building stone and for lining steel furnaces.
4. It forms a protective coating on the surface which prevents further action on the metal below the surface.
5. Calcium is prepared by electrolyzing fused calcium chloride. Sodium is prepared by electrolyzing fused sodium chloride.
6. To deoxidize copper, to harden lead for cables and storage battery grids, and as an alloy with lead for bearings.
7. Limestone, calcite, marble, and shells.
8. (*a*) Water containing ions such as calcium and magnesium which form precipitates with soap. (*b*) Water which does not contain these ions. (*c*) Temporary hardness is due to HCO_3^- ions which are removed by boiling. Permanent hardness is due to more stable negative ions such as SO_4^{--} ions.
9. By allowing a concentrated solution of common salt to stand overnight in contact with the zeolite—the law of mass action.
10. (*a*) Calcium carbonate. (*b*) Calcium oxide. (*c*) Calcium hydroxide. (*d*) Calcium oxide. (*e*) Calcium hydroxide.
11. Carbon dioxide.
12. Put a drop of hydrochloric acid on the coral. Carbon dioxide will be liberated in the reaction which results.
13. $CaCO_3 + H_2O + CO_2 \rightleftarrows Ca(HCO_3)_2$.
14. Powdered limestone and clay are mixed in the proper proportions and heated strongly in a cement kiln.
15. Red fireworks are made with strontium nitrate; green fireworks are made with barium nitrate.

Group B

16. $Mg^{++}(aq) + 2OH^-(aq) \rightarrow Mg(OH)_2(s)$
    $Mg(OH)_2 + 2HCl \rightarrow MgCl_2 + 2H_2O$
    $Mg^{++} + 2Cl^- \rightarrow Mg + Cl_2(g)$
17. So the magnesium will evaporate readily at the temperature of the reaction.
18. Beryllium compounds are covalent, extensively hydrolyzed in water, and very poisonous.
19. Because the atoms of the members of the Calcium family have two electrons which hold the atoms more compactly in the solid metal.
20. Calcium ion may combine with two molecules of water of hydration, while sodium ion does not.
21. Calcium is a less active metal and the calcium hydroxide produced by the reaction coats the surface of the calcium.
22. (a) Ca^{++} and Mg^{++}. (*b*) HCO_3^-. (*c*) SO_4^{--}.
23. Ion exchange reaction, with the formation of a precipitate.
24. (*a*) $CaSO_4 + Na_2CO_3 \rightarrow CaCO_3(s) + Na_2SO_4$.
    (*b*) $Mg(HCO_3)_2 + Na_2CO_3 \rightarrow MgCO_3(s) + 2NaHCO_3$.
25. As a drying agent, to lay dust, to retard the drying of concrete, and to melt ice and snow.
26. (*a*) By gently heating gypsum. (*b*) By uniting with water and forming crystals.
27. As a filler in making paper and in paints.
28. (*a*) Carbon dioxide. (*b*) For making solutions of salts whose positive ions form insoluble carbonates. For making solutions whose pH must be precisely regulated.
29. Each mole of $Ca(OH)_2$ added supplies 2 moles of OH^- ions and 1 mole of Ca^{++} ions, but neutralizes 2 moles of HCO_3^- ions to form 2 moles of CO_3^{--} ions. Thus, 2 moles of Ca^{++} ions are precipitated, 1 mole from the hard water and 1 mole from the lime water.

Solutions to Problems pp. 500-501

Group A

1. $100\bar{0}$ kg – (0.125 × $100\bar{0}$ kg) = 875 kg $CaCO_3$

$$\begin{array}{ccccc} 875\,\text{kg} & & X & & \\ CaCO_3 & \rightarrow & CaO & + & CO_2 \\ 1\,\text{mole} & & 1\,\text{mole} & & \end{array}$$

1 mole $CaCO_3$ = $10\bar{0}$ g 1 mole CaO = 56.1 g

$$X = \frac{875\text{ kg } CaCO_3}{10\bar{0}\text{ g/mole}} \times \frac{1\text{ mole CaO}}{\text{mole } CaCO_3} \times \frac{56.1\text{ g}}{\text{mole}} = 491\text{ kg CaO}$$

2.

$$\begin{array}{ccccc} 86.0\text{ kg} & & & & X \\ 2CaSO_4 \cdot 2H_2O & \rightarrow & (CaSO_4)_2 \cdot H_2O & + & 3H_2O \\ 2\,\text{moles} & & & & 3\,\text{moles} \end{array}$$

1 mole $CaSO_4 \cdot 2H_2O$ = 172 g 1 mole H_2O = 18.0 g

$$X = \frac{86.0\text{ kg } CaSO_4 \cdot 2H_2O}{172\text{ g/mole}} \times \frac{3\text{ moles } H_2O}{2\text{ moles } CaSO_4 \cdot 2H_2O} \times \frac{18.0\text{ g}}{\text{mole}} = 13.5\text{ kg } H_2O$$

3. $Be_3Al_2Si_6O_{18}$: formula weight, 537, Weight of beryllium, 27. 27/537 = 0.050 or 5.0%
4. 1000 kg × $\frac{24}{40}$ = 600 kg magnesium.

5. 4.0×10^6 kg $MgCl_2 \times \frac{24\text{ g Mg}}{95\text{ g } MgCl_2} = 1.0 \times 10^6$ kg Mg

Group B

6. Dolomite is 24.3/184 Mg = 13.2%. 13.2% × 0.950 = 12.5% magnesium in sample.
7. $100\bar{0}$ kg × 0.810 = $81\bar{0}$ kg

$$\begin{array}{ccccc} 81\bar{0}\text{ kg} & & & & X \\ CaCO_3 & \rightarrow & CaO & + & CO_2(g) \\ 1\,\text{mole} & & & & 1\,\text{mole} \\ 10\bar{0}\text{ g/mole} & & & & 44.0\text{ g/mole} \end{array}$$

$$X = 81\bar{0}\text{ kg } CaCO_3 \times \frac{\text{mole}}{10\bar{0}\text{ g}} \times \frac{1\text{ mole } CO_2}{\text{mole } CaCO_3} \times \frac{44.0\text{ g}}{\text{mole}}$$

$X = 356$ kg CO_2

8. 356 kg $CO_2 \times \frac{10^3\text{ g}}{\text{kg}} \times \frac{\text{liter}}{1.98\text{ g}} = 1.80 \times 10^5$ liters CO_2

9. 1.80×10^5 liters $CO_2 \times \frac{760\text{ mm}}{72\bar{0}\text{ mm}} \times \frac{293^\circ\text{ K}}{273^\circ\text{ K}} = 2.04 \times 10^5$ liters CO_2

10. $pV = nRT$

$T = 273^\circ + 28^\circ = 301^\circ$ K

$$p = \frac{95\bar{0}\text{ mm}}{760\text{ mm/atm}} = 1.25\text{ atm}$$

$V = 50\bar{0}\text{ m}^3 \times 10^3$ liter/m^3 = 5.00×10^5 liter

$R = 8.21 \times 10^{-2}$ liter atm/mole °K

$$n = \frac{pV}{RT} = \frac{1.25 \text{ atm} \times 5.00 \times 10^5 \text{ liter}}{8.21 \times 10^{-2} \text{ liter atm/mole °K} \times 301° \text{ K}}$$

$n = 2.53 \times 10^4$ mole

Chapter 25
Answers to Questions pp. 536-537

Group A

1. The metallic character of transition elements is attributed to the small number of electrons in the outermost shell of their atoms.
2. The basic-oxygen process, because it makes large batches rapidly and under carefully controlled conditions.
3. The blue is due to an added dye. Iron(II) tannate is colorless but oxidizes on exposure to air to form black iron(III) tannate.
4. All are found as the native metal. They are soft, dense, and inactive metals. They all have an oxidation state of +1 in some compounds.
5. It contains 14/24 gold.
6. (*a*) Crude copper before it is refined electrolytically. (*b*) Because escaping gases form blisters as the copper cools from the liquid state.
7. The ore is crushed and rocky impurities are washed away. The impure metal is then melted with a flux to get rid of rocky impurities that were not washed away.
8. (*a*) No. (*b*) Silver sulfide.
9. In moist air the coating is a basic copper carbonate. Sulfur dioxide in the air may also form basic copper sulfate. Probably both are formed in most cases. Both are green compounds.
10. For coating metal instead of zinc, for making negative plates of nickel-cadmium storage battery, and for control rods of nuclear reactors.
11. The zinc sulfide is roasted, converting the sulfide to zinc oxide. The oxide is then reduced with carbon to zinc metal. The zinc is purified by distillation.
12. An alloy of mercury with one or more other metals is known as amalgam.
13. The mercury(I) ion consists of two mercury atoms joined by a single covalent bond and each in the +1 oxidation state having given up its remaining outer shell electron, the structure being $(Hg{:}Hg)^{++}$.

Group B

14. The carbon particles form an electrochemical cell with the iron in the presence of moisture. This corrodes the iron.
15. $K_4Fe(CN)_6$ forms the blue precipitate, $KFeFe(CN)_6 \quad H_2O$ with iron(III) ions. $K_3Fe(CN)_6$ forms the same blue precipitate with iron(II) ions.
16. Student's answer should include considerations of speed, quality, and versatility.
17. The metals used in early times were largely the ones that occur in the native state. Those difficult to extract from their ores have been obtained within the last century because of increased knowledge of chemistry which made their extraction possible.
18. The silver can be removed by reacting with concentrated nitric acid or hot, concentrated sulfuric acid.
19. The silver is removed by heating the "mud" with concentrated sulfuric acid. Scrap copper then precipitates silver from the silver sulfate.
20. Copper, zinc, and nickel.
21. By washing the mixture with hot water and filtering. Lead chloride is soluble in hot water.
22. Enough gold and silver are recovered to pay for the process.
23. Finely-powdered ore is leached in a dilute solution of sodium cyanide. The gold forms a soluble sodium aurocyanide. Zinc is used to precipitate the gold from this solution.

24. Crushed silver ore is roasted with salt to form silver chloride. It is then leached in tanks with sodium cyanide solution, forming soluble sodium argenticyanide. Zinc is used to precipitate the silver.
25. Zinc hydroxide is basic in the presence of strong acids, but acts as an acid in the presence of a strong base.

Solutions to Problems p. 537

Group A

1. $165 \text{ g Fe} \times \frac{126.8 \text{ g } FeCl_2}{55.8 \text{ g Fe}} = 375 \text{ g } FeCl_2$

2. $375 \text{ g } FeCl_2 \times \frac{162.3 \text{ g } FeCl_3}{126.8 \text{ g } FeCl_2} = 48\bar{0} \text{ g } FeCl_3$

3. 4Fe = 223. 9O = 144. 6H = 6. Formula weight = 373.
   $223 \div 373 \times 100\% = 59.8\%$

4. $X = \frac{10\bar{0} \text{ g Ag} \times 17\bar{0} \text{ g } AgNO_3}{108 \text{ g Ag}} = 157 \text{ g } AgNO_3$

5. $\frac{112 \text{ g Fe}}{16\bar{0} \text{ g } Fe_2O_3} \times 87.0\% \text{ } Fe_2O_3 = 60.9\% \text{ Fe}$

6. 1.0×10^6 metric tn $\times$ $4.0 \times 10^{-2} = 4.0 \times 10^4$ metric tn.

Group B

7. 8.0×10^4 metric tn $SiO_2 \times \frac{10\bar{0} \text{ g } CaCO_3}{6\bar{0} \text{ g } SiO_2} = 1.3 \times 10^5$ metric tn $CaCO_3$

8. (*a*) 8.7×10^5 metric tn $Fe_2O_3 \times \frac{84 \text{ g CO}}{16\bar{0} \text{ g } Fe_2O_3} = 4.6 \times 10^5$ metric tn CO

   (*b*) 4.6×10^5 metric tn CO $\times \frac{24 \text{ g C}}{56 \text{ g CO}} = 2.0 \times 10^5$ metric tn C.

9. $FeSO_4 + H_2SO_4 + HNO_3 \rightarrow Fe_2(SO_4)_3 + NO + H_2O$

   $\overset{+2}{3Fe} \rightarrow \overset{+3}{3Fe} + 3e^-$

   $\overset{+5}{N} + 3e^- \rightarrow \overset{+2}{N}$

   $6FeSO_4 + 3H_2SO_4 + 2HNO_3 \rightarrow 3Fe_2(SO_4)_3 + 2NO + 4H_2O$

10. $\overset{0}{3Ag} \rightarrow \overset{+1}{3Ag} + 3e^-$

    $\overset{+5}{N} + 3e^- \rightarrow \overset{+2}{N}$

    $3Ag + 4HNO_3 \rightarrow 3AgNO_3 + NO + 2H_2O$

11. $\overset{0}{Cu} \rightarrow \overset{+2}{Cu} + 2e^-$

$\overset{+6}{S} + 2e^- \rightarrow \overset{+4}{S}$

$Cu + 2H_2SO_4 \rightarrow CuSO_4 + SO_2 + 2H_2O$

Chapter 26

Answers to Questions pp. 555-556

Group A

1. Many clays, rocks, and other mineral materials.
2. Silver-white in color; density 2.7 g per cm^3; ductile; malleable; not as tenacious as copper, brass, or steel; good conductor of electricity; can be welded, spun, or cast; difficult to solder.
3. They are oxides of aluminum.
4. Aluminum oxide is an infusible substance. It can be dissociated by being dissolved in fused cryolite.
5. To produce a carbon-free metal. Also because aluminum is a more powerful reducing agent.
6. $KAl(SO_4)_2 \cdot 12H_2O$; $NaAl(SO_4)_2 \cdot 12H_2O$; $KFe(SO_4)_2 \cdot 12H_2O$; $KCr(SO_4)_2 \cdot 12H_2O$.
7. The salt brines of Searles Lake, California.
8. It dissolves metallic oxides leaving a clean metallic surface.
9. Flint, jasper, chalcedony, sand, carnelian, onyx, and agate.
10. (*a*) Second among the elements of the earth's crust. (*b*) The abundance of its compounds, such as sand, sandstone, quartz, and silicate rocks.
11. (*a*) Sand and coke. (*b*) $SiO_2 + 3C \rightarrow SiC + 2CO$.
12. Silicates of sodium and calcium.
13. Sand, limestone, and sodium carbonate.
14. (*a*) It is a sodium aluminum borosilicate glass. (*b*) Its coefficient of expansion is only one-third that of ordinary soda-lime glass.
15. As a by-product.
16. Arsenic is a deadly poison.

Group B

17. (*a*) $2Al + 6HCl \rightarrow 2AlCl_3 + 3H_2$. (*b*) $2Al + 2NaOH + 2H_2O \rightarrow 2NaAlO_2 + 3H_2$.
18. Cathode: $Al^{+++} + 3e^- \rightarrow Al$. Anode: $2O^{--} \rightarrow O_2 + 4e^-$.
19. $Al(OH)_3 + NaOH \rightarrow NaAlO_2 + 2H_2O$.
    $Al(OH)_3 + 3HCl \rightarrow AlCl_3 + 3H_2O$.
20. Aluminum cannot be produced economically from clay.
21. Availability of cheap electricity usually generated by waterpower.
22. (*a*) Small atomic radius; (*b*) ionization energy intermediate between metals and non-metals; (*c*) electronegativity of about 2; (*d*) form covalent bonds; (*e*) low electric conductivity at ordinary temperature; electric conductivity increases with an increase in temperature.
23. Boric acid and alcohol are placed in a porcelain dish. The alcohol is ignited. The edges of the flame have a green color when boric acid is present.
24. (*a*) Quartz crystals consist of tetrahedra with a central silicon atom bonded to four oxygen atoms at the corners. Each oxygen atom serves as the corner of two such tetrahedra. (*b*) If quartz is to be broken, the strong silicon-oxygen bonds must be broken. (*c*) Fused quartz does not have the regular structure of quartz crystals. The arrangement of the tetrahedra is more random.
25. (*a*) A compound containing a chain of alternate silicon and oxygen atoms, with organic groups attached to the silicon atoms. (*b*) As lubricants, as a varnish for the windings of electric motors, as a water repellent, and in automobile and furniture polishes.
26. Silicon carbide has a structure similar to diamond, but with alternate carbon atoms replaced by silicon atoms.

27. Small amounts of mineral impurities discolor the glass.
28. By using small amounts of metallic compounds. Cobalt makes glass blue; chromium, green; iron, pale-green; manganese, amethyst; and colloidal gold, red.
29. Because it has a very low coefficient of expansion.
30. Cobalt.
31. Yes. The similarity in crystal structure gives them similar hardness.
32. (*a*) As molecules. (*b*) It may be safely used as an antiseptic and eye wash.
33. Because it does not react with water directly to form silicic acid.
34. Because of the hydrocarbon groups attached to the silicon atoms.

Solutions to Problems pp. 556-557

Group A

1. $$\begin{array}{ccccccc} X & & Y & & & & 10.0\text{ kg} \\ 2Al & + & Fe_2O_3 & \rightarrow & Al_2O_3 & + & 2Fe \\ 2\text{ moles} & & 1\text{ mole} & & & & 2\text{ moles} \end{array}$$

$1\text{ mole Al} = 27.0\text{ g} \qquad 1\text{ mole } Fe_2O_3 = 160\text{ g} \qquad 1\text{ mole Fe} = 55.8\text{ g}$

$$X = \frac{10.0\text{ kg Fe}}{55.8\text{ g/mole}} \times \frac{2\text{ moles Al}}{2\text{ moles Fe}} \times \frac{27.0\text{ g}}{\text{mole}} = 4.84\text{ kg Al}$$

$$Y = \frac{10.0\text{ kg Fe}}{55.8\text{ g/mole}} \times \frac{1\text{ mole } Fe_2O_3}{2\text{ moles Fe}} \times \frac{160\text{ g}}{\text{mole}} = 14.3\text{ kg } Fe_2O_3$$

2. $27.0 \div 458 \times 100\% = 5.90\%$ Al

3. $$\begin{array}{ccccccccc} 50.0\text{ g} & & 100\text{ g} & & & & & & X \\ 2Al & + & 2NaOH & + & 2H_2O & \rightarrow & 2NaAlO_2 & + & 3H_2 \\ 2\text{ moles} & & 2\text{ moles} & & & & & & 3\text{ moles} \\ 27.0\text{ g/mole} & & 40.0\text{ g/mole} & & & & & & 22.4\text{ liters/mole} \end{array}$$

By inspection: NaOH is in excess.

$$X = \frac{50.0\text{ g Al}}{27.0\text{ g/mole}} \times \frac{3\text{ moles } H_2}{2\text{ moles Al}} \times \frac{22.4\text{ liters}}{\text{mole}} = 62.3\text{ liters } H_2$$

4. $Ca_2B_6O_{11} \cdot 5H_2O = 412$, $6B = 66$
$66 \div 412 \times 100\% = 16\%$ Boron

5. $$\begin{array}{ccccccc} 1.000\text{ kg} & & & & X & & \\ Na_2CO_3 & + & SiO_2 & \rightarrow & Na_2SiO_3 & + & CO_2 \\ 1\text{ mole} & & & & 1\text{ mole} & & \end{array}$$

$1\text{ mole } Na_2CO_3 = 106\text{ g} \qquad 1\text{ mole } Na_2SiO_3 = 122\text{ g}$

$$X = \frac{1.000\text{ kg } Na_2CO_3}{106\text{ g/mole}} \times \frac{1\text{ mole } Na_2SiO_3}{\text{mole } Na_2CO_3} \times \frac{122\text{ g}}{\text{mole}} = 1.15\text{ kg } Na_2SiO_3$$

6. $$\begin{array}{ccccccc} 1.000\text{ kg} & & & & & & X \\ Na_2CO_3 & + & SiO_2 & \rightarrow & Na_2SiO_3 & + & CO_2 \\ 1\text{ mole} & & & & & & 1\text{ mole} \end{array}$$

$1\text{ mole } Na_2CO_3 = 106\text{ g} \qquad 1\text{ mole } CO_2 = 22.4\text{ liters}$

$$X = \frac{100\text{ g } Na_2CO_3}{106\text{ g/mole}} \times \frac{1\text{ mole } CO_2}{\text{mole } Na_2CO_3} \times \frac{22.4\text{ liters}}{\text{mole}} = 211\text{ liters } CO_2$$

7. $\underset{2\text{ moles}}{2Sb} + 3Cl_2 \rightarrow \underset{2\text{ moles}}{2SbCl_3}$ (10.0 g Sb; X)

1 mole Sb = 122 g 1 mole $SbCl_3$ = 228 g

$$X = \frac{10.0\text{ g Sb}}{122\text{ g/mole}} \times \frac{1\text{ mole } SbCl_3}{\text{mole Sb}} \times \frac{228\text{ g}}{\text{mole}} = 18.7\text{ g } SbCl_3$$

Group B

8. $$\text{As : H} = \frac{96.2\text{ g}}{74.9\text{ g/mole}} : \frac{3.85\text{ g}}{1.01\text{ g/mole}} = \frac{1.28\text{ mole}}{1.28\text{ mole}} : \frac{3.81\text{ mole}}{1.28\text{ mole}}$$

As : H = 1.00 : 2.98 or 1.00 : 3.00
Empirical formula = AsH_3 formula weight = 77.9
g-mol wt = D × molar vol = 3.48 g/liter × 22.4 liters = 78.0 g
molecular formula = AsH_3

9. $\underset{1\text{ mole}}{Na_2B_4O_7} + 5H_2O + H_2SO_4 \rightarrow \underset{4\text{ moles}}{4H_3BO_3} + Na_2SO_4$ (5.00; X)

1 mole $Na_2B_4O_7$ = 202 g 1 mole H_3BO_3 = 61.8 g

$$X = 5.00\text{ lb } Na_2B_4O_7 \times \frac{454\text{ g}}{1\text{ lb}} \times \frac{1\text{ mole}}{202\text{ g}} \times \frac{4\text{ moles } H_3BO_3}{\text{mole } Na_2B_4O_7} \times \frac{61.8\text{ g}}{\text{mole}} = 2768.64\text{ g } H_3BO_3$$

10. $\underset{1\text{ mole}}{SiO_2} + \underset{3\text{ moles}}{3C} \rightarrow \underset{1\text{ mole}}{SiC} + 2CO$ (X; Y; 1.00 tn)

1 mole SiO_2 = 60.1 g 1 mole C = 12.0 g 1 mole SiC = 40.1 g

$$X = \frac{1.00\text{ tn SiC}}{40.1\text{ g/mole}} \times \frac{1\text{ mole } SiO_2}{\text{mole SiC}} \times \frac{60.1\text{ g}}{\text{mole}} = 1.50\text{ tn } SiO_2$$

$$Y = \frac{1.00\text{ tn SiC}}{40.1\text{ g/mole}} \times \frac{3\text{ moles C}}{\text{mole SiC}} \times \frac{12.0\text{ g}}{\text{mole}} = 0.896\text{ tn C}$$

11.

Te-122:	0.0246 mole	× 121.903 g/mole	=	3.00 g
Te-123:	0.0087 mole	× 122.904 g/mole	=	1.1 g
Te-124:	0.0461 mole	× 123.903 g/mole	=	5.71 g
Te-125:	0.0699 mole	× 124.904 g/mole	=	8.73 g
Te-126:	0.1871 mole	× 125.903 g/mole	=	23.56 g
Te-128:	0.3179 mole	× 127.905 g/mole	=	40.66 g
Te-130:	0.3448 mole	× 129.907 g/mole	=	44.79 g
				127.6 g = g-at wt
I-127:	1.0000 mole	× 126.904 g/mole	=	126.90 g = g-at wt

at wt Te = 127.6
at wt I = 126.9

Chapter 27
Answers to Questions pp. 570-571

Group A

1. (*a*) In the air. (*b*) In plant and animal proteins and in sodium nitrate and potassium nitrate.
2. (*a*) Daniel Rutherford. (*b*) By removing carbon dioxide from air which had been exhausted by the respiration of small animals.
3. (*a*) Liquid oxygen. (*b*) The separation of liquid nitrogen from liquid oxygen by fractional distillation.
4. (*a*) $NH_4Cl(aq) + NaNO_2(aq) \rightarrow N_2(g) + 2H_2O(l) + NaCl(aq)$.
   (*b*) An exchange reaction with the decomposition of one of the products.
5. Lower burning magnesium into the bottle of gas. If the magnesium continues to burn and forms a white solid which on treatment with water releases ammonia, the original gas in the bottle was nitrogen.
6. The process of converting elementary nitrogen into nitrogen compounds.
7. (*a*) Growing crops which restore nitrogen compounds to the soil, and electric storms.
   (*b*) Manufacturing ammonia and manufacturing calcium cyanamid.
8. (*a*) To shift the equilibrium to the right. (*b*) To increase the speed of the reaction.
9. (*a*) A pyramid with the hydrogen atoms forming an equilateral triangular base and the nitrogen atom forming the apex. (*b*) sp^3-*s* hybrid bonds with an unshared electron pair at the top of the nitrogen atom.
10. Because it is doubtful if there are any molecules of ammonium hydroxide in an ammonia-water solution.
11. Because it emulsifies the film of grease, yet leaves no streaky residue as it dries.
12. Nitric acid vapors would destroy rubber quickly.
13. The nitric acid produced must be driven off by heating.
14. $4NH_3 + 5O_2 \rightarrow 4NO + 6H_2O$. $2NO + O_2 \rightarrow 2NO_2$. $3NO_2 + H_2O \rightarrow 2HNO_3 + NO$.
15. Sunlight promotes its decomposition, forming nitrogen dioxide in sufficient amounts to color the acid yellow.

Group B

16. (*a*) The gas molecules acquire more energy as work is done on them in crowding them together. (*b*) The molecules lose energy as they do work in spreading apart.
17. The strong triple covalent bond between the atoms of the molecule.
18. As a blanketing atmosphere during the processing of food—inactivity. Used with argon in filling electric lamps—inactivity, pressure reduces filament evaporation. To make ammonia—reacts with hydrogen at moderate temperatures, high pressures, and in the presence of a catalyst. Etc.
19. It must be alkaline.
20. Corn removes nitrogen compounds from the soil; lima beans replace them.
21. (*a*) $Ca(OH)_2(s) + 2NH_4NO_3(s) \rightarrow 2NH_3(g) + 2H_2O(g) + Ca(NO_3)_2(s)$. (*b*) $OH^- + NH_4^+ \rightarrow NH_3(g) + H_2O(g)$. (*c*) They are all the same.
22. Ammonia molecules are polar, and the forces of attraction between them are stronger.
23. Ammonia molecules are polar, methane molecules are not. Water is a polar solvent. The solubility follows the general rule that like dissolves like.
24. Ammonia will dissolve and dissociate many salts and form complex ions with metallic ions.
25. The zinc will reduce the nitrogen of the nitric acid more readily than it will reduce the hydrogen.
26. The colorless nitrogen monoxide when released to the air is oxidized to reddish-brown nitrogen dioxide.
27. (*a*) −3. (*b*) −2. (*c*) 0. (*d*) +3. (*e*) +5.
28. Add iron(II) sulfate solution to the unknown and pour concentrated sulfuric acid down the side to the bottom of the test tube. A nitrate will cause a brown layer of nitrosyl iron(II) sulfate to form at the junction of the acid and the other mixture.

Solutions to Problems p. 572

Group A

1. $\underset{1\text{ mole}}{CaCN_2} + 3H_2O \rightarrow CaCO_3 + \underset{2\text{ moles}}{2NH_3}$ (160 g $CaCN_2$; X)

1 mole $CaCN_2$ = 80.1 g 1 mole NH_3 = 17.0 g

$$X = 160\ g\ CaCN_2 \times \frac{\text{mole}}{80.1\ g} \times \frac{2\text{ moles } NH_3}{1\text{ mole } CaCN_2} \times 17.0\ g/\text{mole} = 68.0\ g\ NH_3$$

2. $\underset{1\text{ mole}}{N_2(g)} + \underset{3\text{ moles}}{3H_2(g)} \rightarrow \underset{2\text{ moles}}{2NH_3(g)}$ (X; Y; 200 liters)

By inspection: X = 100 liters nitrogen; Y = 300 liters hydrogen

3. $\underset{1\text{ mole}}{NaNO_3} + H_2SO_4 \rightarrow NaHSO_4 + \underset{1\text{ mole}}{HNO_3}$ (X; 15 g)

1 mole $NaNO_3$ = 85.0 g 1 mole HNO_3 = 63.0 g

$$X = 15\ g\ HNO_3 \times \frac{\text{mole}}{63.0\ g} \times \frac{1\text{ mole } NaNO_3}{1\text{ mole } HNO_3} \times 85.0\ g/\text{mole} = 20\ g\ NaNO_3$$

Group B

4. $\underset{1\text{ mole}}{NH_4Cl} + \underset{1\text{ mole}}{NaNO_2} \rightarrow \underset{1\text{ mole}}{N_2(g)} + 2H_2O + NaCl$ (10 g NH_4Cl; 10 g $NaNO_2$; X)

1 mole NH_4Cl = 53.5 g 1 mole $NaNO_2$ = 69.0 g 1 mole N_2 = 22.4 liters
The NH_4Cl is in excess.

$$X = 10\ g\ NaNO_2 \times \frac{\text{mole}}{69.0\ g} \times \frac{1\text{ mole } N_2}{1\text{ mole } NaNO_2} \times 22.4\ \text{liters/mole} = 3.2\ \text{liters } N_2$$

5. $\underset{1\text{ mole}}{KNO_3} + H_2SO_4 \rightarrow KHSO_4 + \underset{1\text{ mole}}{HNO_3}$ (50.0 g (80.0%); X)

1 mole KNO_3 = 101.1 g 1 mole HNO_3 = 63.0 g

$$X = 50.0\ g\ KNO_3 \times 0.800 \times \frac{\text{mole}}{101.1\ g} \times \frac{1\text{ mole } HNO_3}{1\text{ mole } KNO_3} \times 63.0\ g/\text{mole} = 24.9\ g\ HNO_3$$

6. $\underset{3\text{ moles}}{3Cu} + 8HNO_3 \rightarrow \underset{3\text{ moles}}{3Cu(NO_3)_2} + \underset{2\text{ moles}}{2NO(g)} + 4H_2O$ (254 g Cu; X; Y)

1 mole Cu = 63.5 g 1 mole $Cu(NO_3)_2$ = 187.5 g

$$X = 254\ g\ Cu \times \frac{\text{mole}}{63.5\ g} \times \frac{3\text{ moles } Cu(NO_3)_2}{3\text{ moles Cu}} \times 187.5\ g/\text{mole} = 750\ g\ Cu(NO_3)_2$$

(*b*) 1 mole NO = 22.4 liters

$$Y = 254\ g\ Cu \times \frac{\text{mole}}{63.5\ g} \times \frac{2\text{ moles NO}}{3\text{ moles Cu}} \times 22.4\ \text{liters/mole} = 59.7\ \text{liters NO}$$

Chapter 28
Answers to Questions pp. 586–588

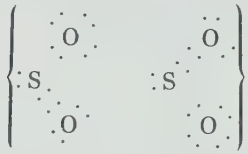
Group A

1. Between 500 and 2000 feet underground in Texas and Louisiana near the Gulf of Mexico.
2. (*a*) To melt the sulfur. (*b*) To force the melted sulfur and water mixture to the surface. (*c*) There is usually a layer of quicksand between the surface of the ground and the sulfur bed.
3. (*a*) Odorless. (*b*) The odor of decayed eggs. (*c*) A suffocating, choking odor.
4. This was the result of a change in the sulfur from amorphous sulfur back to rhombic sulfur.
5. For making sulfur dioxide, carbon disulfide, sulfuric acid, and other sulfur compounds. For making matches, fireworks, gunpowder, sulfur dyes, agricultural sprays, and in vulcanization of rubber.
6. (*a*) Fe_2S_3. (*b*) As_2S_5. (*c*) CuS. (*d*) HgS. (*e*) Ag_2S. (*f*) SnS_2. (*g*) Sb_2S_3.
7. The decay of proteins in which sulfur was present and the burning of coal in which sulfur was present as an impurity.
8. (*a*) $HgCl_2(aq) + H_2S(g) \rightarrow HgS(s) + 2HCl(aq)$
   $PbCl_2(aq) + H_2S(g) \rightarrow PbS(s) + 2HCl(aq)$
   $2SbCl_3(aq) + 3H_2S(g) \rightarrow Sb_2S_3(s) + 6HCl(aq)$
   (*b*) $Hg^{++}(aq) + S^{--}(aq) \rightarrow HgS(s)$
   $Pb^{++}(aq) + S^{--}(aq) \rightarrow PbS(s)$
   $2Sb^{+++}(aq) + 3S^{--}(aq) \rightarrow Sb_2S_3(s)$
9. Copper, zinc, lead, silver, nickel, arsenic, antimony.
10. (*a*) In the manufacture of varnishes, matches, and viscose rayon. (*b*) The solvent action of carbon disulfide is involved in the manufacture of varnishes and matches. Its chemical properties are involved in the production of viscose rayon.
11. From volcanic gases or mineral waters, as a product of the combustion of coal, and from roasting sulfide ores.
12. (*a*) $S(s) + O_2(g) \rightarrow SO_2(g)$
    (*b*) $Cu(s) + 2H_2SO_4(aq) \rightarrow CuSO_4(aq) + 2H_2O(l) + SO_2(g)$
13. (*a*) Upward displacement of air. (*b*) Its solubility in water and greater density than air.
14. Making sulfuric acid.
15. (*a*) $H_2SO_3 + NaOH \rightarrow NaHSO_3 + H_2O$. (*b*) $H_2SO_3 + 2NaOH \rightarrow Na_2SO_3 + 2H_2O$.
16. Because the mixture of sulfur dioxide and oxygen comes in contact with the catalyst.
17. By slowly adding sulfuric acid to water with constant stirring.
18. To remove the oxide coating from iron or steel before the metal is plated.
19. Sulfuric acid acts as a dehydrating agent to absorb the water formed in the reaction between the nitric acid and the glycerol, thus causing the reaction to proceed more readily.
20. Copper(II) sulfate, for copper plating; iron(II) sulfate, for water purification; zinc sulfate, for making lithopone; calcium sulfate, in gypsum; etc.
21. (*a*) Add hydrochloric acid to the unknown salt solution in a test tube and warm gently. Expose moist lead(II) acetate paper to the fumes at the mouth of the test tube. If the paper turns brownish-black, a sulfide was present. (*b*) Add barium chloride solution to the salt solution. If a sulfate was present, a white precipitate, insoluble in concentrated hydrochloric acid, will form.
22. To prevent the formation of interfering white precipitates of barium oxalate, barium carbonate, barium sulfite, or barium phosphate.

Group B

23. (*a*) S_8. (*b*) It needlessly complicates an equation by making it necessary to multiply all the coefficients by 8.
24. At the melting point sulfur is a pale-yellow easily flowing liquid. As it is heated further it becomes more viscous until above about 160° C it hardly pours. The color at this time has changed from light yellow to reddish-brown to almost black. Near the boiling point the fluidity increases again.
25. Endothermic. Monoclinic sulfur is stable only at a higher temperature than rhombic.

26. (*a*) $FeS(s) + H_2SO_4(aq) \rightarrow FeSO_4(aq) + H_2S(g)$
(*b*) Exchange reaction.
(*c*) $FeS(s) + 2H_3O^+(aq) \rightarrow H_2S(g) + 2H_2O(l) + Fe^{++}(aq)$
27. If there is an abundance of oxygen available, sulfur dioxide and water vapor are the products. If an equal volume of oxygen is available, the products are principally water vapor, sulfur dioxide, and sulfur.
28. $H_2S + H_2O_2 \rightarrow 2H_2O + S(g)$



30. Easy to liquefy. It requires only three atmospheres pressure at room temperature to liquefy it.
31. It combines readily with water to form sulfurous acid.
32. (*a*) $Hg + 2H_2SO_4 \rightarrow HgSO_4 + SO_2(g) + 2H_2O$. (*b*) $2Cu_2S + 3O_2 \rightarrow 2Cu_2O + 2SO_2(g)$.
33. Solutions of hydrogen sulfites have a lower pH than solutions of sulfites. Except for the sulfites of Group I metals, most other sulfites are insoluble while the hydrogen sulfites are soluble.
34. For the exothermic oxidation of sulfur dioxide the temperatures must be kept at 400° C. The heat of this oxidation reaction is used to preheat the entering reactants, thus regulating the temperature in the catalyzer.
35. Sulfur trioxide does not dissolve readily in pure water, but does dissolve readily in 98% sulfuric acid forming pyrosulfuric acid.
36. It has a high temperature and also is a powerful oxidizing agent.

Answers to Problems pp. 588–589

Group A

1. $\underset{1\text{ mole}}{S} + O_2 \rightarrow \underset{1\text{ mole}}{SO_2}$ (1.0 kg S; X)

$1\text{ mole S} = 32.1\text{ g} \qquad 1\text{ mole } SO_2 = 64.1\text{ g}$

$$X = 1.0\text{ kg S} \times \frac{\text{mole}}{32.1\text{ g}} \times \frac{1\text{ mole } SO_2}{1\text{ mole S}} \times 64.1\text{ g/mole} = 2.0\text{ kg } SO_2$$

2. (*a*) $\underset{2\text{ moles}}{2H_2S(g)} + \underset{3\text{ moles}}{3O_2(g)} \rightarrow 2H_2O + 2SO_2$ (5.0 liters; X)

$$X = 5.0\text{ liters } H_2S \times \frac{3\text{ moles } O_2}{2\text{ moles } H_2S} = 7.5\text{ liters } O_2$$

(*b*) $$7.5\text{ liters } O_2 \times \frac{100\%\text{ air}}{21\%\ O_2} = 36\text{ liters air}$$

3. Formula weight of PbS = 207.2 + 32.1 = 239.3.

$$\text{Pb: } \frac{(207.2)\text{ Pb}}{(239.3)\text{ PbS}} \times 100\%\text{ PbS} = 86.6\%\text{ Pb}$$

$$\text{S: } \frac{(32.1)\text{ S}}{(239.2)\text{ PbS}} \times 100\%\text{ PbS} = 13.4\%\text{ S}$$

4. $\underset{1\text{ mole}}{CS_2} + 3O_2(g) \rightarrow \underset{1\text{ mole}}{CO_2(g)} + 2SO_2$ (39.0 g CS_2; X)

1 mole CS_2 = 76.2 g 1 mole CO_2 = 22.4 liters

$$X = 39.0\text{ g } CS_2 \times \frac{\text{mole}}{76.2\text{ g}} \times \frac{1\text{ mole } CO_2}{1\text{ mole } CS_2} \times 22.4\text{ liters/mole} = 11.5\text{ liters } CO_2$$

5. Formula weight of H_2SO_4 = 2(1.0) + 32.1 + 4(16.0) = 98.1

$$\text{H: } \frac{(2.0)\text{ H}}{(98.1)\ H_2SO_4} \times 100\%\ H_2SO_4 = 2.0\%\text{ H}$$

$$\text{S: } \frac{(32.1)\text{ S}}{(98.1)\ H_2SO_4} \times 100\%\ H_2SO_4 = 32.7\%\text{ S}$$

$$\text{O: } \frac{(64.0)\text{ O}}{(98.1)\ H_2SO_4} \times 100\%\ H_2SO_4 = 65.2\%\text{ O}$$

6. $\underset{1\text{ mole}}{Na_2SO_3} + H_2SO_4 \rightarrow H_2O + \underset{1\text{ mole}}{SO_2(g)} + Na_2SO_4$ (X; 1.00 liter)

1 mole Na_2SO_3 = 126.1 g 1 mole SO_2 = 22.4 liters

$$X = 1.00\text{ liter } SO_2 \times \frac{\text{mole}}{22.4\text{ liters}} \times \frac{1\text{ mole } Na_2SO_3}{1\text{ mole } SO_2} \times 126.1\text{ g/mole} = 5.63\text{ g } Na_2SO_3$$

7. For each mole of ZnS roasted, one mole of H_2SO_4 may be theoretically produced.
1 mole ZnS = 97.5 g 1 mole H_2SO_4 = 98.1 g

$$50\bar{0}\text{ tons ZnS} \times \frac{10^3\text{ kg}}{\text{ton}} \times \frac{98.1\text{ g } H_2SO_4}{97.5\text{ g ZnS}} = 5.03 \times 10^5\text{ kg } H_2SO_4$$

Group B

8. 1 mole H_2S = 34.1 g

$$50\bar{0}\text{ ml} \times \frac{273^\circ\text{ C}}{30\bar{0}^\circ\text{ C}} \times \frac{74\bar{0}\text{ mm}}{76\bar{0}\text{ mm}} \times \frac{1\text{ liter}}{10^3\text{ ml}} \times \frac{34.1\text{ g}}{22.4\text{ liters}} = 0.674\text{ g}$$

9. From 1 mole of sulfur, one mole of sulfuric acid may be theoretically produced.
1 mole S = 32.1 g 1 mole H_2SO_4 = 98.1 g

$$5.00\text{ kg S} \times 0.995 \times \frac{98.1\text{ g } H_2SO_4}{32.1\text{ g S}} = 15.2\text{ kg } H_2SO_4$$

10. $\underset{4\text{ moles}}{4FeS_2} + 11O_2 \rightarrow 2Fe_2O_3 + \underset{8\text{ moles}}{8SO_2(g)}$ ($120\bar{0}$ kg; X)

1 mole FeS_2 = 120.0 g 1 mole SO_2 = 22.4 liters

$$X = 120\bar{0}\text{ kg } FeS_2 \times \frac{10^3\text{ g}}{\text{kg}} \times \frac{1\text{ mole}}{120.0\text{ g}} \times \frac{8\text{ moles } SO_2}{4\text{ moles } FeS_2} \times 22.4\text{ liters/mole}$$

$$\times \frac{298^\circ\ K}{273^\circ\ K} \times \frac{760\ mm}{740\ mm} = 5.02 \times 10^5 \text{ liters } SO_2$$

11. (*a*) $\underset{1\text{ mole}}{Fe} + \underset{1\text{ mole}}{H_2SO_4} \rightarrow \underset{1\text{ mole}}{FeSO_4} + H_2$

$1\text{ mole Fe} = 55.8\text{ g} \qquad 1\text{ mole } FeSO_4 = 151.9\text{ g}$

$$X = 140\text{ kg Fe} \times \frac{\text{mole}}{55.8\text{ g}} \times \frac{1\text{ mole } H_2SO_4}{1\text{ mole Fe}} \times 151.9\text{ g/mole} = 381\text{ kg } FeSO_4$$

(*b*) $1\text{ mole } H_2SO_4 = 98.1\text{ g}$

$$Y = 140\text{ kg Fe} \times \frac{\text{mole}}{55.8\text{ g}} \times \frac{1\text{ mole } H_2SO_4}{1\text{ mole Fe}} \times 98.1\text{ g/mole} \times \frac{100\%\text{ acid}}{95\%\ H_2SO_4} = 259\text{ kg acid}$$

12. $\underset{1\text{ mole}}{Cu} + \underset{2\text{ moles}}{2H_2SO_4} \rightarrow CuSO_4 + 2H_2O + \underset{1\text{ mole}}{SO_2(g)}$

$1\text{ mole Cu} = 63.5\text{ g} \qquad 1\text{ mole } H_2SO_4 = 98.1\text{ g} \qquad 1\text{ mole } SO_2 = 22.4\text{ liters}$
The copper is in excess.

$$X = 100\text{ g } H_2SO_4 \times \frac{\text{mole}}{98.1\text{ g}} \times \frac{1\text{ mole SO}}{2\text{ moles } H_2SO_4} \times 22.4\text{ liters/mole} = 11.4\text{ liters } SO_2$$

13. $$[S^{--}] = \frac{6.8 \times 10^{-24}}{[H_3O^+]^2}$$

pH	$[H_3O^+]$	$[S^{--}]$
2	10^{-2}	6.8×10^{-20}
4	10^{-4}	6.8×10^{-16}
6	10^{-6}	6.8×10^{-12}
8	10^{-8}	6.8×10^{-8}
10	10^{-10}	6.8×10^{-4}

14. (*a*) $[Hg^{++}][S^{--}] = 3.5 \times 10^{-52}$ $\qquad$ $[Fe^{++}][S^{--}] = 4.9 \times 10^{-18}$
$10^{-2}\ [S^{--}] = 3.5 \times 10^{-52}$ $\qquad$ $10^{-2}\ [S^{--}] = 4.9 \times 10^{-18}$
$[S^{--}] = 3.5 \times 10^{-50}$ $\qquad$ $[S^{--}] = 4.9 \times 10^{-16}$

If $[S^{--}]$ is approximately 10^{-17} HgS will precipitate but FeS will not.

$$[S^{--}] = \frac{6.8 \times 10^{-24}}{[H_3O^+]^2}$$

$$[H_3O^+]^2 = \frac{6.8 \times 10^{-24}}{[S^{--}]} = \frac{6.8 \times 10^{-24}}{10^{-17}} = 6.8 \times 10^{-7}$$

$[H_3O^+] = 10^{-3}$ pH = 3 or lower will effect separation.

(*b*) $[Bi^{+++}]^2[S^{--}]^3 = 6.8 \times 10^{-97}$ $\qquad$ $[Zn^{++}][S^{--}] = 1.1 \times 10^{-21}$
$(10^{-2})^2[S^{--}]^3 = 6.8 \times 10^{-97}$ $\qquad$ $10^{-2}\ [S^{--}] = 1.1 \times 10^{-21}$
$[S^{--}]^3 = 6.8 \times 10^{-93}$ $\qquad$ $[S^{--}] = 1.1 \times 10^{-19}$
$[S^{--}] = 1.9 \times 10^{-31}$

If $[S^{--}]$ is approximately 10^{-20} Bi_2S_3 will precipitate but ZnS will not.

$$[S^{--}] = \frac{6.8 \times 10^{-24}}{[H_3O^+]^2}$$

$$[H_3O^+]^2 = \frac{6.8 \times 10^{-24}}{[S^{--}]} = \frac{6.8 \times 10^{-24}}{10^{-20}} = 6.8 \times 10^{-4}$$

$[H_3O^+] = 10^{-2}$ pH = 2 or lower will effect separation

(*c*) $[Co^{++}][S^{--}] = 8.7 \times 10^{-23}$ $[Ni^{++}][S^{--}] = 1.8 \times 10^{-21}$
$10^{-2}\ [S^{--}] = 8.7 \times 10^{-23}$ $10^{-2}\ [S^{--}] = 1.8 \times 10^{-21}$
$[S^{--}] = 8.7 \times 10^{-21}$ $[S^{--}] = 1.8 \times 10^{-19}$
These two ions cannot be separated by sulfide precipitation.

15. $[Ba^{++}][SO_4^{--}] = 1.5 \times 10^{-9}$ $10^{-2}[SO_4^{--}] = 1.5 \times 10^{-9}$
$[SO_4^{--}] = 1.5 \times 10^{-7}$
16. $[Pb^{++}][S^{--}] = 8.4 \times 10^{-28}$ $10^{-2}\ [S^{--}] = 8.4 \times 10^{-28}$
$[S^{--}] = 8.4 \times 10^{-26}$

Chapter 29
Answers to Questions pp. 605-606

Group A

1. They are too strongly electronegative.
2. Astatine, iodine, bromine, chlorine, fluorine.
3. Salt producer.
4. (*a*) A special carbon steel container. (*b*) A wax or plastic lined container.
5. $Xe + F_2 \rightarrow XeF_2$. $XeF_2 + F_2 \rightarrow XeF_4$. $XeF_4 + F_2 \rightarrow XeF_6$.
6. As a catalyst in the manufacture of high-octane gasoline, in the manufacture of synthetic cryolite, and for etching glass.
7. (*a*) Sodium chloride. (*b*) Sodium.
8. $\overset{+4\ -2}{MnO_2} + 4\overset{+1-1}{HCl} \rightarrow \overset{+2\ -1}{MnCl_2} + 2\overset{+1-2}{H_2O} + \overset{0}{Cl_2}$
Chlorine is oxidized; manganese is reduced.
9. When inhaled in small quantities, it produces about the same symptoms as a bad head cold. In larger quantities, it causes death.
10. (*a*) More dense than air and very soluble in water. (*b*) Upward displacement of air.
11. Add silver nitrate and nitric acid. A white precipitate of silver chloride indicates the presence of a chloride.
12. For use in manufacturing anti-knock fluids for gasoline.
13. (*a*) $2Br^-(aq) + Cl_2(g) \rightarrow 2Cl^-(aq) + Br_2(l)$. (*b*) Displacement reaction.
14. Dark-red volatile liquid, three times as dense as water, and moderately soluble in water. Readily soluble in carbon tetrachloride, carbon disulfide, and in water solutions of bromides.
15. Natural underground brines.
16. Some of the solvent, alcohol, may have evaporated, causing the solution to increase in concentration to a point where it will damage skin tissue.
17. (*a*) For thyroxine which controls the rate at which the body uses food energy.
(*b*) Iodized salt and sea food.

Group B

18. There are no elements more electronegative than fluorine.
19. $\overset{+1\ -2}{2H_2O} + \overset{+2-1}{2XeF_2} \rightarrow \overset{0}{2Xe} + \overset{0}{O_2} + \overset{+1\ -1}{2H_2F_2}$
20. Odorless, nonflammable, nontoxic, easily condensed to a liquid by pressure at ordinary temperatures.
21. Because hydrogen fluoride molecules are quite polar.
22. Because otherwise they would react with one another.
23. (*a*) Hydrogen. (*b*) When a paraffin candle burns in chlorine, the chlorine unites with the hydrogen of the paraffin and leaves the carbon uncombined.
24. (*a*) The yellow-green color of chlorine is imparted to the chlorine water. (*b*) The chlorine unites with water and forms hypochlorous acid and hydrochloric acid which are colorless.

25. Oxygen.
26. Only in the sense that it is a proton donor. Liquid hydrogen chloride contains no hydronium ions.
27. Add carbon tetrachloride and chlorine water. Shake. If the carbon tetrachloride layer is colored red, bromide ions were present.
28. (*a*) Grayish-black. (*b*) Dark-brown. (*c*) Violet. (*d*) Violet.
29. (*a*) HF, HCl, HBr, HI. (*b*) The polarity of the bonds should decrease as one goes from HF to HI. (*c*) $HF(g) + H_2O(l) \rightarrow H_2O^+(aq) + F^-(aq)$. $HCl(g) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)$. $HBr(g) + H_2O(l) \rightarrow H_3O^+(aq) + Br^-(aq)$. $HI(g) + H_2O(l) \rightarrow H_3O^+(aq) + I^-(aq)$.
30. (*a*) HF: Concentration of molecules much higher than the concentrations of hydronium and fluoride ions. HCl, HBr, and HI: Concentrations of hydronium and halide ions are much higher than the concentration of hydrogen halide molecules. (*b*) Hydrogen fluoride molecules are more stable than hydronium and fluoride ions. Hydrogen chloride, hydrogen bromide, and hydrogen iodide molecules are less stable than hydronium ions and the corresponding halide ions.
31. Aluminum is not as highly electropositive a metal as sodium and calcium.

Answers to Problems p. 606

Group A

1. $$2NaCl + 2H_2O \xrightarrow{\text{elect}} \underset{2\text{ moles}}{\overset{X}{2NaOH}} + H_2 + \underset{1\text{ mole}}{\overset{71\bar{0}\text{ kg}}{Cl_2}}$$

   1 mole Cl_2 = 71.0 g 1 mole NaOH = 40.0 g

   $$X = 71\bar{0}\text{ kg } Cl_2 \times \frac{\text{mole}}{71.0\text{ g}} \times \frac{2\text{ moles NaOH}}{1\text{ mole } Cl_2} \times 40.0\text{ g/mole} = 80\bar{0}\text{ kg NaOH}$$

2. $$\underset{2\text{ moles}}{\overset{X}{2NaBr}} + MnO_2 + 2H_2SO_4 \rightarrow \underset{1\text{ mole}}{\overset{10.0\text{ g}}{Br_2}} + Na_2SO_4 + MnSO_4 + 2H_2O$$

   1 mole NaBr = 102.9 g 1 mole Br_2 = 159.8 g

   $$X = 10.0\text{ g } Br_2 \times \frac{\text{mole}}{159.8\text{ g}} \times \frac{2\text{ moles NaBr}}{1\text{ mole } Br_2} \times 102.9\text{ g/mole} = 12.9\text{ g NaBr}$$

3. (*a*) $$\underset{1\text{ mole}}{\overset{Y}{Cl_2}} + \underset{1\text{ mole}}{\overset{X}{Ca(OH)_2}} \rightarrow \underset{1\text{ mole}}{\overset{25\bar{0}\text{ g}}{Ca(ClO)Cl}} + H_2O$$

   1 mole $Ca(OH)_2$ = 74.1 g 1 mole Ca(ClO)Cl = 127.1 g

   $$X = 25\bar{0}\text{ g } Ca(ClO)Cl \times \frac{\text{mole}}{127.1\text{ g}} \times \frac{1\text{ mole } Ca(OH)_2}{1\text{ mole } Ca(ClO)Cl} \times 74.1\text{ g/mole} = 146\text{g } Ca(OH)_2$$

   (*b*) 1 mole Cl_2 = 71.0 g

   $$Y = 25\bar{0}\text{ g } Ca(ClO)Cl \times \frac{\text{mole}}{127.1\text{ g}} \times \frac{1\text{ mole } Cl_2}{1\text{ mole } Ca(ClO)Cl} \times 71.0\text{ g/mole} = 14\bar{0}\text{ g } Cl_2$$

4. $$Zn + \underset{1\text{ mole}}{\overset{11.2\text{ liters}}{Cl_2(g)}} \rightarrow \underset{1\text{ mole}}{\overset{X}{ZnCl_2}}$$

   1 mole Cl_2 = 22.4 liters 1 mole $ZnCl_2$ = 136.4 g

   $$X = 11.2\text{ liters } Cl_2 \times \frac{\text{mole}}{22.4\text{ liters}} \times \frac{1\text{ mole } ZnCl_2}{1\text{ mole } Cl_2} \times 136.4\text{ g/mole} = 68.2\text{ g } ZnCl_2$$

Group B

5. $\underset{\text{1 mole}}{\underset{390\text{ g}}{CaF_2}} + \underset{\text{1 mole}}{\underset{600\text{ g(95\%)}}{H_2SO_4}} \rightarrow CaSO_4 + \underset{\text{2 moles}}{\underset{X}{2HF}}$

1 mole CaF_2 = 78.1 g 1 mole H_2SO_4 = 98.1 g 1 mole HF = 20.0 g
H_2SO_4 is in excess.

$$X = 390\text{ g } CaF_2 \times \frac{\text{mole}}{78.1\text{ g}} \times \frac{2\text{ moles HF}}{1\text{ mole } CaF_2} \times 20.0\text{ g/mole} = 200\text{ g HF}$$

6. Formula weight of $C_2H_4Br_2$ = 2(12.0) + 4(1.0) + 2(79.9) = 187.8

$$\text{Br: } \frac{159.8\text{ (Br)}}{187.8\text{ } (C_2H_4Br_2)} \times 100\%\ C_2H_4Br_2 = 85.1\%\text{ Br}$$

7. $\underset{\text{2 moles}}{\underset{468\text{ g}}{2\,NaCl}} \xrightarrow{\text{elect}} 2\,Na + \underset{\text{1 mole}}{\underset{X}{Cl_2}}$

1 mole NaCl = 58.5 g 1 mole Cl_2 = 22.4 liters

$$X = 468\text{ g NaCl} \times \frac{\text{mole}}{58.5\text{ g}} \times \frac{1\text{ mole } Cl_2}{2\text{ moles NaCl}} \times 22.4\text{ liters/mole} = 89.6\text{ liters } Cl_2$$

8. $$5 \times 250\text{ ml} \times \frac{1\text{ liter}}{10^3\text{ ml}} \times \frac{273^\circ\text{ C}}{300^\circ\text{ C}} \times \frac{750\text{ mm}}{760\text{ mm}} = 1.12\text{ liters } Cl_2 \text{ at STP.}$$

$$\underset{\text{4 moles}}{\underset{X}{4HCl}} + \underset{\text{1 mole}}{\underset{Y}{MnO_2}} \rightarrow MnCl_2 + 2H_2O + \underset{\text{1 mole}}{\underset{\text{1.12 liters}}{Cl_2(g)}}$$

1 mole HCl = 36.5 g 1 mole Cl_2 = 22.4 liters

$$X = 1.12\text{ liters } Cl_2 \times \frac{\text{mole}}{22.4\text{ liters}} \times \frac{4\text{ moles HCl}}{1\text{ mole } Cl_2} \times 36.5\text{ g/mole} \times \frac{100\%\text{ acid}}{38\%\text{ HCl}} \times \frac{1\text{ ml}}{1.20\text{ g}} =$$

16.0 ml cone HCL
1 mole MnO_2 = 86.9 g

$$Y = 1.12\text{ liters } Cl_2 \times \frac{\text{mole}}{22.4\text{ liters}} \times \frac{1\text{ mole } MnO_2}{1\text{ mole } Cl_2} \times 86.9\text{ g/mole} = 4.34\text{ g } MnO_2$$

Chapter 30

Answers to Questions pp. 626–627

Group A

1. (*a*) Henri Becquerel. (*b*) By accident, Becquerel found that uranium ore gives off invisible rays which affect a photographic plate in the same way light does.
2. Pitchblende was four times as radioactive as the amount of uranium it contained warranted.
3. Radium is more than 1,000,000 times as radioactive as uranium.
4. In producing luminous paint.
5. The length of time required for the decay of one-half of any number of atoms of a radioactive nuclide.
6. The nucleus.
7. The nuclide is changed to one having an atomic number two less and a mass number four less.
8. The nuclide is changed to one having an atomic number one greater and a mass number remaining the same.

9. By measuring the amount of parent nuclide and daughter nuclides in the sample.
10. Radioactive decay, nuclear disintegration, fission, and fusion.
11. The emission of an alpha particle decreases the mass of the nucleus; it also decreases the number of protons and neutrons by an even number. Beta particle emission brings the neutron-proton ratio nearer the value for stable nuclei of the same mass number.
12. By Chadwick during the bombardment of beryllium by alpha particles.
13. Because they do not carry any charge. They are not attracted or repelled by a nucleus, but can easily penetrate it.
14. It may go through the atom, or it may be caught by the nucleus. It may cause the nucleus to disintegrate.
15. In medical treatments and as tracers in medical and industrial research.
16. (*a*) $^{238}_{92}U$, $^{235}_{92}U$, and $^{234}_{92}U$. (*b*) 99.3%, 0.7%, and insignificant traces, respectively.
17. (*a*) Fission is the break-up of a very heavy nucleus of an atom into medium-weight parts. (*b*) It is produced in $^{235}_{92}U$ by the capture of a slow neutron, which produces a highly unstable nucleus. The nucleus immediately undergoes fission.
18. The amount of radioactive material required to provide sufficient neutrons to sustain a chain reaction.
19. Because of the heat generated during the chain reaction. If this heat were allowed to accumulate it would soon destroy the reactor.
20. (*a*) A fusion reaction in which four hydrogen nucleu combine and form a helium nucleus. (*b*) A thermonuclear bomb.

Group B

21. Because it is more important for its radioactivity than for its chemical or physical properties.
22. Radioactivity is a property of the nuclei of atoms. Nuclei are not altered during chemical combination.
23. The particles given off by the radiation can easily penetrate the black paper to affect the film.
24. The more radioactive a substance is, the faster it will discharge an electroscope.
25. Beyond bismuth in the periodic table.
26. (*a*) Four times that of a protium atom; helium ions; near one-tenth the speed of light. (*b*) 1/1837 as heavy as a protium atom; electrons; near the speed of light.
27. High energy electromagnetic waves.
28. $^{226}_{88}Ra \rightarrow {}^{222}_{86}Rn + {}^{4}_{2}He$.
29. $^{214}_{82}Pb \rightarrow {}^{214}_{83}Bi + {}^{0}_{-1}e$.
30. $^{214}_{84}Po \rightarrow {}^{210}_{82}Pb + {}^{4}_{2}He$.
    $^{210}_{82}Pb \rightarrow {}^{210}_{83}Bi + {}^{0}_{-1}e$.
31. (*a*) The lightest and heaviest elements. (*b*) The intermediate-weight elements. (*c*) The greater the binding energy per nuclear particle is, the more stable the nucleus is.
32. Mass, ratio of neutrons to protons, and the even-odd nature of the number of neutrons, and the number of protons.
33. Each type of nuclear reaction results in more stable nuclei, because the products are nuclei with greater binding energies per nuclear particle.
34. (*a*) Rutherford. (*b*) $^{14}_{7}N + {}^{4}_{2}He \rightarrow {}^{17}_{8}O + {}^{1}_{1}H$.
35. By bombarding lithium with high-speed protons and comparing the actual loss of mass with Einstein's predicted loss.
36. (*a*) It is a spiral path. (*b*) The combined action of the magnetic field and the alternating voltage on the dees.
37. $^{238}_{92}U$ is bombarded with neutrons to produce the unstable isotope of uranium, $^{239}_{92}U$. This emits a beta particle to form neptunium, $^{239}_{93}Np$. This in turn emits a beta particle to form plutonium, $^{239}_{94}Pu$.
38. (*a*) By bombarding normally nonradioactive substances with neutrons or deuterons. (*b*) $^{59}_{27}Co + {}^{1}_{0}n \rightarrow {}^{60}_{27}Co$. (*c*) Used in treatment of cancer.

39. (*a*) A chain reaction is one in which one of the starting materials is also one of the products. (*b*) The fission of $^{235}_{92}U$ is started by a neutron. During the fission process, several neutrons are emitted. Some of these initiate fission in other $^{235}_{92}U$ nuclei, which in turn produce other neutrons.
40. It is constructed of blocks of natural uranium alternated with blocks of graphite which serve as a moderator. Control rods and cooling apparatus are incorporated also.

Part IV

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WAS-513 Breeder Reactors
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Part V

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Teacher's Guide for High School Regents Chemistry, Curriculum Report no. 422-1500-SS, Bureau of Curriculum Development, Board of Education, The City of New York, 101 Livingston St., Brooklyn, N.Y. 11201, 1968.

Lab Bench Experiments in Chemistry, Reprint Department *Chemistry*, 1970, 1155 Sixteenth St., N.W., Washington 20036.

Physics Handbook, New York State Education Department, Albany, N.Y. 12224, 1970 (The 1956 edition is also available).

Sutton, R. M., *Demonstration Experiments in Physics*, McGraw-Hill, 1938.

Sources of Equipment: Aloe Scientific Div. of Brunswick Corp., 1150 Flower St., Los Angeles, Cal. 90015.

Atomic Accessories, Inc., 811 West Merrick Rd., Valley Stream, N.Y. 11580.

Atomic Laboratories, Inc., 3088 Claremont Ave., Berkeley, Cal. 94705.

Baird-Atomic, 35 University Road, P.O. Box 400, Cambridge, Mass., 02138.

Cambosco (Scientific Co.), 342 Western Ave., Boston, Mass. 02135.

Cenco (Central Scientific Co.), 2600 S. Kostner Ave., Chicago, Ill. 60623.

Edmund Scientific Co., 103 Gloucester Pike, Barrington, N.J. 08007.

Educational Materials and Equipment Co., Box 63, Bronxville, N.Y. 10708.

Fisher Scientific Co., 633 Greenwich St., New York, N.Y. 10014.

Klinger Scientific Apparatus Corp., 83-45 Parsons Blvd., Jamaica, New York 11432.

Labasco (Standard Scientific Supply Corp.), 808 Broadway, New York, N.Y. 10003.

La Pine Scientific Co., 6001 S. Knox Ave., Chicago, Ill. 60629.

Macalaster Scientific Corp., Route 111 and Everett Turnpike, Nashua, N.H. 03060.

Metrologic Instruments Inc., 143 Harding Ave., Bellmar, N.J. 08030.

Nuclear-Chicago Corp., 301 E. Howard Ave., Des Plaines, Ill. 60618.

**Note:* In this Teacher Edition, references to Alyea and Sutton, *Tested Demonstrations in Chemistry*, appear simply as Alyea.

REAC (Radio Equipment and Accessories Corp.), 665 Merrick Rd., Lynbrook, N.Y.

Sargent-Welch Scientific Co., 7300 N. Linder Ave., Skokie, Ill. 60076.

Scientific Glass Apparatus Co., 737 Broad St., Bloomfield, N.J. 07003.

Sesco, Inc., 1312 South Thirteenth St., Vincennes, Indiana 47591.

Stansi Scientific Co., 1231-45 N. Honore St., Chicago, Ill. 60622.

Wilkens-Anderson Co., 4515 W. Division St., Chicago, Ill. 60651.

Will Corporation, Wilco Drive, Rochester, N.Y. 14603.

**Note:* Many of the companies listed above supply models in great variety. The following companies specialize in model production.

Dyna-Slide Co., 1566 Sherman Ave., Evanston, Ill. 60201 (Stereofold).

Ealing Corp., 2225 Massachusetts Ave., Cambridge, Mass. 02140. (Ask for Atomic Model Reference Library.) (SRM Molecular, Orbital and Crystal Model Systems, Scale Model Molecular Models.)

Edmund Scientific Co., Barrington, N.J. 08007 (Atomic, Molecular, and Crystal Models).

Klinger Scientific Co., 83145 Parsons Blvd., Jamaica, N.Y. 11432.

Macro Models, 948 Howard St., San Francisco, Cal. 94103.

Scientific Educational Products Corp., 30 East 42nd St., N.Y., N.Y. 10017.

Science Related Materials, Inc., P.O. Box 1422, Janesville, Wi. 53545 (Large-scale models for class display student kits).

Thomas, Arthur H., Co., Nine St. at Third, P.O. Box 779, Philadelphia, Pa.

Class-constructed Models: Student-constructed may be assembled from plastic sphere, eggs, and teardrops which may be purchased from such companies as:

Ace Plastic Co., 91-30 Van Wyck Expressway, Jamaica, N.Y. 11435.

Plasteel Corp., 26970 Princeton St., Inkster, Mich.

Star Band Co., Inc., Board and Commerce Sts., Portsmouth, Va.

Student Kits: Cenco, or O. H. Jons Glas Co., Ltd., 219 Broadview Ave., Toronto, Canada (Jonsglas Student Molecular Models.)

Sargent-Welch: 5736 Student Molecular Kit. (See list of suppliers for address.)

Printed matter related to models:

Gordon, A. J., "A Survey of Atomic and Molecular Models," J. Chem. Educ., 47, 30 (1970). Lists model types, commercially available model sets, suppliers, and specific references to articles.

Wells, A. F., *Models in Structural Inorganic Chemistry*, Oxford University Press, New York, N.Y. (1970), paper. Kit available.

Sanderson, R. T., *Principles and Construction of New Chemistry Aids*, Bureau of Audio-Visual Instruction, State University of Iowa, Iowa City, Ia.

King, L. C., "Molecular Architecture," *Chemistry*, 37, 12 (1964); "Molecular Architecture," *The Science Teacher*, 33, 6 (1966).

Bassow, H., *Construction and Use of Atomic and Molecular Models*, Pergamon Press Inc., 395 Sawmill River Road, Elmsford, N.Y. 10523.

Journals

Chemistry, American Chemical Society, 1155 Sixteenth St., N.W., Washington, D.C. 20036.

Journal of Chemical Education, Chemical Education Publishing Co., 20th & Northampton Sts., Easton, Pa. 18042.

Scientific American, Scientific American Inc., 415 Madison Avenue, New York, N.Y. 10017.

The Physics Teacher, American Institute of Physics, 335 East 45th Street, New York, N.Y. 10017.

The Science Teacher, National Science Teachers Association, 1201 Sixteenth St., Washington, D.C. 20036.

Science World, Scholastic Magazines Inc., 902 Sylvan Ave., Englewood Cliffs, New Jersey 07632.

Indexes, Guides, Handbooks:

American Chemical Society, *Aids and Sources for the Science Teacher: A Guide for the Science Teacher.*

Chemistry, List of Reprints (100+) Reprint Department, Chemistry, 1155 Sixteenth St., N.W., Washington, D.C. 20036.

Journal of Chemical Education, 25-year Cumulative Index (Vol. 1-25, 1924-49; 10-year Cumulative Index (Vol. 26-35, 1949-58); 10-year Cumulative Index (Vol. 36-45, 1959-68).

Scientific American Offprints, Annual Catalog listing all available reprints, W. H. Freeman Co., 660 Market Street, San Francisco, Cal. 94104.

The Merck Index, Merck & Co., Inc., Rahway, New Jersey 07065.

Handbook of Chemistry and Physics, Annual, Chemical Rubber Co., 18901 Cranwood Pkwy., Cleveland, Ohio.

Note: 1. The April issues of *Journal of Chemical Education* contain a section listing the paperback publications of the National Bureau of Standards Reference Data System, U.S. Dept. of Commerce, National Bureau of Standards, Washington, D.C. 20234.

2. The December issues of such publications as *Chemistry* and *Journal of Chemical Education* include a yearly Index.

3. The September issue of the *Journal of Chemical Education* includes a section, Book buyers' Guide: Texts, Laboratory Manuals, Programmed Instruction, Paperbacks, and Audio-Visual Materials.

Key to Publishers:

Academic Press, Inc., 111 Fifth Ave., New York, N.Y. 10003.
Addison-Wesley Publishing Co., Reading, Mass. 01867.
Allyn & Bacon, Inc., 470 Atlantic Ave., Boston, Mass. 02210.
American Chemical Society, 1155 Sixteenth St., N.W., Washington, D.C. 20036.
Appleton-Century Crofts, Inc., 440 Park Ave. S., New York, N.Y. 10016.
W. A. Benjamin, Inc., South Street, Reading, Mass.
Cambridge University Press, 32 East 57th St., New York, N.Y. 10022.
Chemical Rubber Co., 18901 Cranwood Pkwy., Cleveland, Ohio 44128.
Chemical Education Publishing Co., 20 and Northampton, Easton, Pa. 18042.
Cornell University Press, 124 Roberts Place, Ithaca, N.Y. 14850.
Thomas Y. Corwell Co., 666 Fifth Ave., New York, N.Y. 10019.
Doubleday & Company Inc., 277 Park Ave., New York 10017.
W. H. Freeman & Co., 660 Market St., San Francisco, Cal. 94104.
Harcourt Brace Jovanovich, Inc., 757 Third Ave., New York, N.Y. 10017.
Harper & Row, 10 East 53 St., New York, N.Y. 10022.
D. C. Heath, 125 Sprint St., Lexington, Mass. 02173.
Holt, Rinehart and Winston, Inc., 383 Madison Ave., New York, N.Y. 10017.
Houghton Mifflin Co., 1 Beacon St., Boston, Mass. 02108.
McGraw-Hill Book Co., 1221 Ave. of the Americas, New York, N.Y. 10020.
Macmillan Inc., 866 Third Ave., New York, N.Y. 10022.
Oxford University Press, Inc., 200 Madison Ave., New York, N.Y. 10016.
Prentice-Hall, Inc., Englewood Cliffs, N.J. 07632.
W. B. Saunders Co., 218 West Washington Sq., Philadelphia, Pa. 19105.
U. S. Dept. of Commerce, National Bureau of Standards, Clearing House, Springfield, Va. 22151.
Van Nostrand Reinhold Co., 450 West 33 St., New York, N.Y. 10001.
John Wiley & Sons, Inc., 605 Third Ave., New York, N.Y. 10016.

Audio-visual Aids:

These include films, loops, transparencies—overhead and 35 mm slides, and filmstrips. Identifying Key letters appear in the Key to Suppliers of Audio-visual Aids which follows. The CHEM films and some PSSC films as well as a rather extensive listing in each chapter provide for a wide variety of choice. There are thirteen CHEM STUDY LOOPS available. Ealing loops may now be purchased through Holt, Rinehart and Winston. Loops from many other sources again provide for a wide variety of choice. Holt, Rinehart and Winston provides a set of 100 transparencies which reproduce many of the figures in *Modern Chemistry*. Transparencies from other sources are also suggested. The Visual Products Division, 3M Company has an excellent booklet, "Getting the Most out of your Overhead Projection System," It suggests methods of making your own visuals. Another valuable booklet by Keuffel & Esser Co. is "Products for Overhead Projection."

Note: 1. A set of 120 two-color slides for Masterton and Slowinski's, *Chemical Principles* (a college text) are available from W. B. Saunders Co., West Washington Square, Philadelphia, Pa. 19105.

2. The Visual Produtts Division, 3M Company, 3M Center, St. Paul, Minn. 55101. Has two series for CHEM study: Overhead Projection Transparencies 15-2660-7, 15-2661-5, which can only be purchased in single sets or as a combination.

3. W. B. Saunders Co., W. Washington Sq., Philadelphia, Pa. 19105, has a series for college chemistry (long films, c, s), which comes in either S-8 cartridge, S-8 reel, or 16mm reel. Typical titles, identified as S include: *Acids and Bases in Titration, Estimation of Molecular Size and Avogadro Number, Measurement of Vapor Pressure + The Variation of Solubility with Temperature.*

4. American Chemical Society, 1155 16th St., Washington, D.C. 20036 has a series of audiotape cassettes plus reference manuals. Typical titles include: *Xenon Compounds, Homogeneous Catalytic Reactions of Molecular Hydrogen, Radiocarbon Dating.*

5. John Wiley & Sons, Inc., 605 Third Ave., New York, N.Y. 10016 has a series for college chemistry, Single Concept 8mm Films in Chemical Education, (Kodak or Technicolor cartridge) which includes: *Atomic and Bonding Orbitals, The Sartorius Balance 2400, The Mercury Beating Heart, Corrosion III Aluminum, Model of the Kinetic-Molecular Concept, The Critical Point, the pH Meter, On Writing Equations-I, Reaction Rates, On Writing Equations-II, Energy Cycles.* These are labeled W as loops.

6. GE Films, General Electric Educational Films, Building 705, Corporations Park, Scotia, N.Y. 12302.

Key to Suppliers of Films, Filmstrips, Film Loops, Overhead Transparencies, and Slides:

AEC Audio-Visual Branch, Division of Public Information, U.S. Atomic Energy Commission, Washington, D.C. 20545.
AFL Anargyros Film Library, 1813 Fairburn Ave., Los Angeles, Cal. 90025.

AIM Association Instructional Materials, Association Films, 347 Madison Ave. (Dept. DC), New York, N.Y. 10017.

C Cambosco Scientific Co., Inc., 342 Western Ave., Boston, Mass. 02135. (Overhead transparencies).

CHUR Churchill Films, 622 N. Robertson Blvd., Los Angeles, Cal. 90069.

CORF Coronet Films, 65 E. South Water St., Chicago, Ill. 60601.

DG Denoyer-Geppert Audio-Visuals (Time Mirror), 5235 Ravenswood Ave., Chicago, Ill. 60640. (Filmstrips, overhead transparencies).

DUART Du Art Film Laboratories, Inc., U.S. Government Film Service, 245 W. 55 St., New York, N.Y. 10019.

E Ealing Film-Loops, 2225 Massachusetts Ave., Cambridge, Mass. 02140 (Many may be ordered from Holt, Rinehart and Winston and are identified as H-E or E).

EBF Encyclopaedia Britannica Films, Inc., 425 N. Michigan Ave., Chicago, Ill. 60611, Yale Chemistry Film Series listed as YF.

EYE Eye Gate House, Inc., 146-01 Archer Ave., Jamaica, N.Y. 11435.

FAC Film Associates of California, 11559 Santa Monica Blvd., Los Angeles, Cal. 90025.

HM Houghton Mifflin Co., 110 Tremont St., Boston, Mass. 02107 (Overhead transparencies).

HR Harper & Row, 49 East 33 St., New York, N.Y. 10016 (Overhead transparencies).

HRW Holt, Rinehart and Winston, Inc., 383 Madison Ave., New York, N.Y. 10017 (Transparencies for Modern Chemistry; Ealing loops identified as H-E; transparencies for Project Physics identified as PP).

ICF International Communications Films, 1371 Reynolds Ave., Santa Ana, Cal. 92705.

IFB International Film Bureau Inc., 332 S. Michigan Ave., Chicago, Ill. 60604.

INDU Indiana University, Audio-Visual Center, Bloomington, Indiana 47401.

K & E Keuffel & Esser Co., Hoboken, N.J. 07030. (Overhead transparencies).

LBF LB Films, 3435 Grant Street, Corvallis, Oregon 97330.

LGC Longmans, Green & Co., 48 Grosvernor St., London, W.1., England.

McW Macalaster Scientific Co., Route 111 and Everett Turnpike, Nashua, N.H. 03060 (Walker overhead transparency sets).

MH McGraw-Hill Book Co., Text-Film Division, 1221 Ave. of the Americas, New York, N.Y. 10020 (films, overhead transparencies).

MLA Modern Learning Aids, 1212 Ave. of the Americas, New York, N.Y. 10036 (CHEM films and loops, PSSC films). Also Modern Learning Aids Division (Ward's Natural Science Establishment, Inc.), P.O. Box 302, Rochester, N.Y. 14603.

PPG PPG Industries, Inc., One Gateway Center, Pittsburgh, Pa. 15222.

ROB Peter M. Robeck & Co., Inc., 230 Park Ave., New York, N.Y. 10017.

STER Sterling Movies, Inc., 43 West 61 St., New York, N.Y. 10023.

SUTH Sutherland Educational Films, Inc., 201 N. Occidental Blvd., Los Angeles, Cal. 90026.

W Western Publishing Co., Inc., 850 Third Ave., New York, N.Y. 10022 (35 mm color slides).

Part VI

Word Lists

Words from the HARRIS and JACOBSON List

Chemical words from the Harris and Jacobson grade school, list that most students would know before enrolling for a college preparatory or introductory college chemistry course. Words such as acid and base on this list are also included on the technical vocabulary list because their common or grade school usage is expanded into a more technical meaning during the chemistry course.

*acid
alcohol
alpha
aluminum
ammonia
ammonium
atmosphere ic
barium base
*base
beryllium
boron
bromine
burner
cadmium
calcium
cancer
carbohydrate
carbonate
carburetor
cellulose
ceramic
chemical
chemistry
chlorinate
chlorine
chlorophyll
circulate ion
classification
cobalt
combustion
compression
concentration
condense ation
contaminate
crystalline
Curie
data
diffusion
dilute
dioxide
discharge
drug
electromagnetic
electron ic
electroscope
evaporation
experimentation
extinguisher
fahrenheit
filament
fission able
foil
fusion
galvanometer
geiger
gelatin
generate or
geologist
geology
gram
graphite
helium
hydrochloric
hypothesis
impurity
infrared
iodine
isotope
krypton
limewater
lithium
magnesium
magnetic
manganese
mercuric
molecule ar
neon
neutron
nitrate
nitrogen
nuclei us
oxide
phenomena
phosphorus
photon
physicist
physics
potassium
protein
proton
radioactive
radium
reactor
researcher
silicon
sodium
spectrum
sulfate
sulfide
sulfuric
test-tube
thorium
ultraviolet
uranium
velocity

List B Technical Vocabulary Words

This list contains 350 words that have been identified by Dr. Jacobson's computer analysis as the technical vocabulary of the text. These words either (1) exceed the 9th grade level and/or (2) their technical use in chemistry is beyond the grade level of their common use.

On this list, variants of the root word are shown separated by a space. For example, emit ted ting means that the words emit, emitted, and emitting were included in the talley. An asterisk (*) after the word indicates that the word is used more than 100 times throughout the text. Words are listed in the first chapter in which they occur. The next two chapters in which the word occurs are listed in parenthesis. If the word occurs in more than these chapters, a plus(+) sign indicates that the word is used in chapters after the last one listed. However, if a word occurs in several chapters scattered throughout the book and experience has shown that slow or average students might forget the word from one use to the next, additional chapters have been listed.

Chapter 1

alchemy
analytical (19, 20)
Avogadro (3, 7, 10)
C (Celsius)* (2, 3, +)
calorie (1)
cancelation (12)
cataly st sis tic (2, 16, +)
Celsius (9)
chlorate (2, 8, 14, 22)
chloride* (2, 3, +)
cm (centimeter) (2,3,+)
decompos e ing ition (2,6,8,+)
density ies (9,10,+)
deviation
dimensional
electrolyte
emit ted ting (3,4,22,30)
equilibrium* (6,11,+)
exponential
g (gram)* (2,3,+)
generalization (6,7,+)
gram (2,3,+)
hardness (11,16,22,24,26)
incandescent (22)
inert (5,13,21)
kcal (2,5,6,19)
kinetic s (9,19)
liter* (9,10,15,20)
magnitude (3,6,+)
milli (prefix) (9,10,+)
ml (milliliter) (9,10,+)
mm (millimeter) (9,10,+)
organic (12,16,+)
phlogiston
polonium (4,26,30)
potential (2,11,+)
residue (10,16,25)
significan t ce (3,7,+)
spectra l (3,4,22)
synthesis ized (2,16,+)
tolerance
transitions (4)
wavelength (4,5,22)
X ray (4,5,+)

Chapter 2

analysis (6,7,+)
antimony (4,5,+)
argon (3,5,+)
arsenic (5,25,+)
bombard ment (9,11,
Bunsen (flame) (11,22)
carbide (11,17)
cyclotron (2,30)
disulfide (carbon) (28,29)
endothermic (6,12,19,+)
entropy (12,19,+)
exothermic (6,12,19,+)
heterogeneous (19,20)
homogeneous (11,12,19,+)
insoluble (8,11,14,+)
malleable (22,24,+)
metallic (5,6,8,11,+)
metalloid al (5,6,22,+)
monoxide (7,10,16,+)
nitric (14,24,+)
nonmetal lic (5,6,+)
platinum (3,12,17+)
plutonium (30)
precipitate* (8,13,+)
reactant* (8,10,+)
rhombic (28)
spontaneous ly (9,12,19,+)
theoretical (4,9,+)

Chapter 3

accelerator (particle) (30)
angstrom
at (atomic) (25)
deflect ion
deuterium/deuterons (11,22,30)
evacuated (5,9)
fluoride (5,6,8,29)
gram-atomic weight (7)
lb (pound)
nuclide (11,30)
protium
spectroscope (4,22)
structural (6,14,+)
tritium

Chapter 4

absorb ed (11,19,+)
bright-line spectrum
electromagnetic (5)
electron-configuration (5)
electron-dot (6,17)
frequency (5,19)
notation
orbitals (5,6,17,+)
quantum
stable/stability (5,6,11+)
sublevel (5,22,25)
xenon (5,6,29)

Chapter 5

affinity ies
halogen (8,13,17,+)
ionization* (6,13,+)
periodic (8,14,+)

Chapter 6

algebraic sum
anion (14,15,+)
bond (7,8,+)
bromide (29)
carbon-hydrogen (12,17,+)
cation (14,15,20,+)
coefficient (8,10,21)
covalent* (7,8,11,+)
diatomic (7,8,10,+)
electronegativity (8,11,+)
empirical (7,8,10,+)
hybrid ization (12,16,+)
hydrazine (27)
hydride (22)
ionic* (8,11,+)
methane (7,9,+)
monatomic (8,10,+)
nonpolar (9,11,+)
oxidation number* (7,8,14,+)

oxidiz e ing ation (18,21,25,+)
peroxide (21,22)
polar ity (11,13)
polyatomic (7,14,21,+)
reducing (agent) (8,16,+)
resonance (16,17)
subscript (7,8,30)
symmetrical (11,12,+)
tetrahedron al (11,12,+)
valence (11,16,22,+)

Chapter 7

acetate (15,20)
crystal lized ing (11,12,+)
gram-formula weight (12,15)
gram-molecular weight (8,10,+)
hydroxide* (8,11,13,+)
tetrachloride (carbon) (9,12,+)

Chapter 8

carbonic acid (11,14,16,+)
iodide (19,20,29)
reversible (12,13,+)
stoichiometr y ic (28)

Chapter 9

barometer
Boyle's (law)
Charles' (law)
displacement (10)
eudiometer
Kelvin
logarithm (15)
STP (standard temperature and pressure) (10,+)

Chapter 10

molar ity (11,15)
octane (17)
vapor ize ization (11,12,16,+)

Chapter 11

amorphous
anhydride
anhydrous (15,25,+)
aq* (aqueous) (14,21,+)
deliquescent (22)
diffraction
diffuse (12,13,+)
disperse ion (colloidal) (22,26)
dispersion interaction (22)
efflorescent ce
electrolysis (16,21,+)
ethanol (12,16)
ether (12,17,+)
ethyl (12,13,16,+)
hexagonal (16,22,26)
hydrat e ion* (12,13,+)
hydronium (13,14,+)
hygroscopic (18,22)
interaction (forces) (12,16,22)
LeChatelier's Principle (15,20,+)
moderator (16,24,30)
monoclinic (28)
naphthalene (12,17)
saturated (12,13,+)
volatile (12,16,+)

Chapter 12

adsorb ption (16)
beaker (13,22)
boiling-point (13)
carbon-carbon bond (16,17)
carbon-oxygen (16,17)
carbonated (16)
colloidal
depression (13)
electrolyte (13,21,+)
ethylene (17,18)
freezing point (13)
miscible
molal (13,+)
molality (13,+)
solub le ility* (14,20,+)
solute* (13,14,+)
solvent* (13,14,+)
sucrose (18,28)
supersaturated

Chapter 13

acetic acid (14,15,+)
dimer (17)
dipole (14,21)
dissociation (14,20)
electrode* (16,21)
electrolytic (21,25)
fluoride (17,29)
halide (17,18,29)
ionize* (14,+)
titration

Chapter 14

*acceptor (proton) (20,27,+)
Brønstead
caustic (22)
conductivity (15,22)
conjugate
corrosion (25,27,29)
diprotic (15)
formic (16)
gelatinous (22,25,+)
hypochlor ite ous (27,29)
inorganic (16,17,+)
litmus (15,20)
methyl (15,17,18)
neutral ize ization (15,18,+)
nomenclature (21)
OH^- (ion) (15,20,+)
oxyacid
phenolphthalein (15)
phosphoric acid (22)
protolysis
reagent (20,28,+)

Chapter 15

buret
equivalence point
normal ity ities
pH* (20,28)

Chapter 16

activated
allotrop e ic (28)
boneblack
dehydrat e ion (17,18,28)
destructive distillation (22)
distil lation (17)
flammable (17,18,+)
methanol (18)
ppm (parts per million)
refin e ing (17,25)
synthetic (17,22,+)

Chapter 17

alcohol
alkadiene
alkane
alkene
alkyl
aromatic
benzene
branched-chain (hydrocarbon)
butadiene
butane
butyl
chloroform (18,29)
dimethyl
ethane
ethene (18)
ethyne (18,19)
fractionating tower
homologous
hydrocarbon (22,26,29)
hydrogenation (18,25)
isomers
iso-octane
neopentane
neoprene
nitro (prefix) (27)
pentane
polymer ize ic ization (22,24,26)
propane
resin (18,24)
substitution (18)
vulcaniz e ation (28)
xylene s

Chapter 18

acetaldehyde
acetone

aldehyde
aldol
carbonyl (group)
carboxylic acid
ester ification
formaldehyde
freon
hydrofluoric (25,26,29)
hydroly sis zing (20,22,+)

Chapter 19

activation (energy)
free-energy
kcal/mole (20)
thermo (chemical) (dynamic) (nuclear)
water-gas

Chapter 20

buffer ed

Chapter 21

anode (22,+)
cathode (22,25,+)
electrochemical
electrolyzing (22,24,+)
half-cell
half-reaction
permanganate

Chapter 22

cesium
cobalt-blue glass
complex
disintegrate (26,30)
ductile (25,26)
francium
half-filled (25,27)
luster (26)
metallurgical (25,27)
phosphine
superoxide

Chapter 23

All of the technical vocabulary words for this chapter have been previously introduced.

Chapter 24

amalgam
deionized
extraction (25,29,30)
filter ate (25,26)
flux (25,26)
gypsum
ion-exchange
quicklime
radioactive ity (29,30)
refractory
zeolite s

Chapter 25

basic-oxygen process
crucible (26,28)
cyanide
ingot
mordant (26)
open-hearth
paramagnetic
silica te (26)
smelt er ing (26,28)

Chapter 26

alloy ed
borax
cryolite (29)
thermite

Chapter 27

cyanamid
fixation (nitrogen)
Haber-process

Chapter 28

sulfite
sulfurous (acid)
viscous/viscosity

Chapter 29

radon
tincture

Chapter 30

half-life
radium
synchrotron
transmutation
transuranium

Part VII

Teacher Notes

TEACHER NOTES CHAPTER 1

Section 1.1. Start with a "spectacular." A mixture of Zn powder and flowers of sulfur on an asbestos sheet form a rising white cloud. Add colorless solutions in a base-acid-base sequence to a tall cylinder having an almost invisible quantity of phenolphthalein solution in the bottom. A "wine-to water-to milk" demonstration, using successive solutions of $KMnO_4$, $NaHSO_3$ or H_2SO_3, then $BaCl_2$ is another alternative. Such demonstrations enliven students' interest and should encourage them to ask questions (although detailed answers and discussion of most questions should be deferred until much later).

Refer to articles on pollution. As examples of the chemist's concern with the structure of materials, exhibit ball-and-stick models of H, H_2, H_2O, ethanol, ammonia, and carbon dioxide. Exhibit products invented and produced by chemists and technologists: silicones, nylon, other synthetic fibers, plastics, and medicinals. Also display containers or samples of household chemicals. To emphasize further the relevance of chemistry, have students discuss reasons for studying chemistry, its importance in daily life, and its importance to various professions, medicine, law, forestry, dentistry, pharmacy and others. Exhibit the latest December issue of *Chemistry* magazine. To show how library research helps scientists, have students locate and read the names of some of the articles listed in the Index under "Pollution" and other significant headings such as "Detergents," and "Lead."

Section 1.2. If possible, exhibit photos of important chemists from the past. Also refer to the biographical sketches in the textbook. Show the filmstrip DG 678 *Chemistry and the Nobel Prize.* Past and present issues of *Chemistry* give references to Nobel Prize winners in the December Index. (Use the indices for *Chemists* under "Biography.") Students may be assigned to give reports on such winners. You may wish to show the loop, AFL S-5, *Life of Albert* Einstein.

To illustrate the conservation of mass burn a piece of paper and ask relevant questions. Counterpoise a candle on a balance and have students note loss in weight as it burns. Balance a piece of steel wool on a platform balance. Burn it and have students note gain in weight (mass). A wad of thin Cu turnings also gives good results on a sensitive balance.

Heat HgO and "discover" the element oxygen. Have students report on Lavoisier and the phlogiston theory (*Foundations of Chemistry*, Chapter 1). To accompany this you may also wish to use the film, YF-216, *Decomposition of Mercuric Oxide* (3 min, c, s) based on a reversible reaction. Also refer students to *Chemistry* (Reprint 80) "Chance Favors the Prepared Mind," and (Dec. Indices) for another series, "Great Moments in Chemistry." Refer to the replacement of the phlogiston theory to show how science uses evidence.

Section 1.3. You may wish to follow up or precede the discussion of scientific methods with a demonstration series. Burn Mg ribbon in air. Heat steel wool and/or a strip of Cu in air. To get more spectacular results, heat the metals (hold in tongs) and then insert into bottles of O_2. An appropriate series of questions should cover the following aspects: observing, generalizing, explaining-theorizing, testing. Have students *predict* the behavior of heated platinum both in air and in oxygen. Follow up by performing the demonstration. You may then show one of these films listed in the overprint of the text (Teachers' Edition) to illustrate the use of scientific methods in industry and/or research.

Sections 1.4 and 1.6. **Exhibits:** A variety of objects and materials to help students realize that matter is recognized by use of the senses. **Demonstration:** Weigh

and then inflate a balloon. Weigh inflated. Consider buoyancy. Balance an evacuated 500-ml flash with one-hole stopper and clamp or valve attachment on a simple balance such as a horn-pan balance. Open the valve and note change in mass, demonstrating that air has mass and takes up space. Refer to astronauts on the moon for the mass-weight distinction. Exhibit and use a spring balance and platform balance. Which is heavier, lead or cork? Weigh a buckshot and then a large piece of cork which weighs more. Weigh cubes or cylinders of the same volume of whatever materials are available (Welch, 1138, 1158). You may wish to have students measure the length of side of a cube (metric ruler) and calculate the volume.

Section 1.7. Phases: Exhibit ice, water, and boiling water, and moisten the chalk board with a small quantity of water. Change positions of a solid in a battery jar. Change the position and/or use a variety of different-shaped vessels for an equal volume of water or colored water. Tie the simple experiments together by the idea that matter occupies space in all phases. Open a bottle of colored gas to show diffusion of gasses. Pour a liquid from one container into another to demonstrate fluidity, or use the diffusion of a colored gas into air. "Magically" extinguish a candle burning in a battery jar by pouring invisible CO_2 into the battery jar.

Section 1.8. Properties: **(a)** Physical properties defined: Each is specific. **Demonstration:** Use a variety of substances pinpointing physical properties, such as H_2S for odor, NO_2 for color, NaCl and clay for solubility, etc. **Question:** How are physical properties determined? **(b)** Chemical properties defined: Each is specific. **Demonstration:** Perform selected experiments of these types: Attempt to burn a square of cardboard and a square of asbestos of same shape, size, thickness. Ignite a small piece of yellow phosphorus on asbestos by touching with a heated glass rod; duplicate with sulfur. Put a piece of Mg and one of Cu in separate test tubes of hydrochloric acid. Add a piece of Na (CARE) and one of Fe to separate test tubes of water. Add a solution of NaCl first to a solution of KNO_3 then to a solution $AgNO_3$. In all instances give the names of substances and have students read name labels. Use questions such as: What change did you see? Did you notice an odor? Was there a change in temperature? This procedure trains powers of observation and stresses importance of noting a change as well as recognizing that no perceptible change has occurred. Therefore students will not *always* expect something to happen. Energy: This section involves difficult physical concepts which in a course in physics would require at least one day for development. The teacher may choose to present the material rather than try to develop it wholly or in part according to the suggested plan **(a)** Chemical energy: **Question:** Why is the chemist so interested in whether or not things burn? What accompanies the process of burning? Discuss heat resulting from burning in connection with fuels: home, auto, rocket, etc. **Demonstration:** Heat a stoppered test tube containing a small amount of water until the stopper is shot out. **Question:** How is heat utilized? (Production of motion—mechanical energy.) Where did the heat come from? Explain how the heat caused the stopper to be shot out. **(b)** Mechanical energy. **Demonstration:** Use a single fixed pulley over which two unequal weights are strung. Have the lighter weight on the floor and the other on the table top. Release the weight from the table top. **Question:** What is done when a weight is lifted? (This question should lead to the concept of work as a force resulting in motion. It may be used in connection with the "cannon" experiment above to give the idea that work is done on the stopper since a force produces motion. It may be necessary for the teacher to explain briefly that work is done whenever something is moved, that is, a force results in motion. It may also be expedient for the teacher to define energy.) Where did the energy involved in raising the weight come from? How much more energy does the heavier weight have? (Discussion may lead to the idea that if the weight fell a greater distance to the floor below, it could do more work.) What kind of energy did the weight have on the table top? On the floor? (Potential.) What kind of energy did it possess as it fell? (Kinetic.) What happened to the potential energy? To the kinetic energy? (Transformation.) How much energy does the smaller weight have? How could this be demonstrated? **Demonstration:** Replace the heavier weight (now on the floor) with one smaller than the weight which was raised and which you now place on the table. Repeat the demonstration by releasing the weight on the table. (Class discussion may arise about losses of energy, as by friction.) **(c)** Other forms of energy: energy transformations. Have the class trace back the light in the classroom. Turn it on). **Question:** What happens to the light and heat energy of the bulb? Where does the energy propelling the Polaris submarines come from? **(d)** Conservation of energy. Some attempt should be made to have the law of conservation of energy (in non-nuclear processes) stated or recognized.

Section 1.11. **(a)** Conservation of mass: Weight a photoflash bulb before and after flashing. **(b)** Balance a flask containing two test tubes, one containing a solution of lead nitrate and the other a solution of potassium chromate. Invert the stoppered flask and replace it on the balance. **Question:** What did you observe? (Appearance?) Mass (Weight?) How do you interpret the results? Tell the students that repeated experiments *always* give this result. Why is this statement (interpretation) called a law? What does a law tell us? **(c)** Conservation of mass and energy. **Question:** Where does the energy resulting in the explosion of an atomic bomb or hydrogen bomb come from? Student

report on Sec. 1.11 in the text. **Question:** If this law is true, where does the energy of ordinary chemical reactions come from?

Sections 1.12 and 1.13. Have students discuss the units in which common materials are purchased and measured (cloth, butter, milk; representative of length, weight, and volume units). Have them attempt to explain how a system of measurement, such as English measurements of length, evolved. Ask students who have been abroad how distances are measured in France, how gasoline is measured in France, Mexico, Canada, etc. Have them attempt to explain why there are different systems. Exhibit or demonstrate the units of measurement as the lesson develops. Provide some opportunity for simple conversions.

Sections 1.14 and 1.15. **Demonstrations:** Take Calsius and Fahrenheit temperatures of melting ice and boiling water. Mix hot and cold water in plastic calorimeter as in Experiment 1. To show heat transfer insert a metal U-shaped yolk into hot and cold water samples in plastic cups. Have a student touch the metal yolk. Check the thermometers in each beaker after a sufficient time lapse to reach thermal equilibrium.

Section 1.16. **Activity:** Have students calculate the calories lost by the hot water and the calories gained by the cold water samples in the previous water mixing demonstration. Define kilocalorie.

Section 1.17. **Activities:** Either by individual student experiment, or teacher demonstration with student aid, having a series of measurements made with devices of different degrees of precision. For example a length may be measured first by meter stick and then by caliper or micrometer, or a mass by using balances of differing degrees of precision. By comparing measurements reported by several students and measurements reported by the same student but with different instruments, students can be led to the idea that the degree of accuracy of any measurement depends on one's ability to use the instrument properly and on the precision of the device. Develop the idea of significant figures on this basis and then procedures for their manipulation. **Laboratory Experiment 2,** "Measurements and Accuracy." **Exercise:** Schedule Exercise 1, *Exercises and Experiments in Chemistry*. Present a duplicate of Fig. 2-4, p. 31 from *Modern Physics*, 1972 and utilize data to analyze precision and accuracy. Refer to pp. 30–32 in *Modern Physics*. 1. Uncertainty in measurements: accuracy-absolute errors (experimental) and relative errors (percentage); precision-deviations (absolute, average, relative). Use the textbook procedures in Sec. 17. Utilize student findings in Procedure 9, Experiment 1, *Laboratory Experiments in Chemistry* or use this Procedure as a classroom demonstration (Table 1.5 in text). 2. Concept of significant figures. From student measurements of length, mass, or volume have them realize that observation leads to measurements with one doubtful figure. Go through the computations in the text, Sec. 18. Have students state the reason for rounding off results of computation. Give examples of how the decimal point placement, bar placement, and zeros play important roles in indicating the number of significant figures represented. Have students answer question 18.

Section 1.18. **Exhibit and demonstrate:** use of triple beam balances of varying sensitivities, i.e., 0.1g and 0.01g. Demonstrate use of a balance of the Mettler or Sartorius type to show degree of precision.

Section 1.19. **Activities:** Use the text illustrations and then apply to Sample Problems and problems 7–16.

Films:

PSSC 0104 *Measuring Short Distances* (20 min, s);

PSSC 0105 *Measurement* (21 min, s);

PSSC 0111 *Elements, Mixtures, Compounds* (33 min, c, s);

PSSC 0313 *Conservation of Energy* (27 min, s);

CHEM 4160 *A Research Problem: Inert (?) Gas Compounds* (19 min, c, s);

EBF 702 *The Scientific Method* (16 mm, 12 min, c, b/w, s);

EBF 2068 *Thermometers and How They Work* (16 mm, 11 min, c, b/w, s);

EBF *Classifying Matter—The Need* (35 min, b/w or c, s);

EBF *Laboratory Techniques and Apparatus—I and II* (each 35 min, b/w or c, s). Excellent for training students for laboratory work;

EBF *The Physical States of Matter* (35 min, b/w or c);

EBF *Weighing—The Analytical Balance* (35 min, b/w or c, s);

EBF *Weighing with the Analytical Balance* (18 min, sd, b/w);

EBF *The Top-loading Balance* (21 min, s, c) and/or *The Modern Balance* (16 min, s, c) both on free loan from Mettler Instrument Corporation, 20 Nassau St., Princeton, N. J. 08540. This same company provides a number of filmstrips and transparencies including some dealing with analytical balance theory and sources of weighing errors;

MH *Temperature and Matter* (15 min, s and b/w or c, s);

MH *The Analytical Balance* (13 min, b/w, s) Fisher Scientific Co.:

CORF *Introduction to Chemistry* (16 mm, 11 min, c, b/w, s);

DA MIS677 *Sources of Air Pollution* (16 mm, 5 min, c, s);

DA 1407-1 *Challenging Careers in Chemistry* (16 mm, 23^{+} min, c, s). Research projects and making new products for industry from agricultural commodities;

GE *Man-made Diamond* (16 mm, 11 min, c, s);

INDU FS-1154 *Water for a Thirsty World* (16 mm, 30 min, b/w, s);

ROB *Science and Foresight* (16 mm, 25 min, b/w, s);
SE *The Modern Chemist—Diamond Synthesis* (16 mm, 13 min, c, s);
AIM *The Yale Chemistry Films*
YF-240 *Properties of Mixtures and Compounds* (4 min, c, s);
YF-201 *Using a Pipette* (4 min, s, c);
YF-202 *Using a Filter* (2 min, s, c);
YF-203 *Quantitative Transfer* (17 min, c, s);
YF-207 *Laboratory Burners* (9 min, c, s);
YF-208 *Air Damped Analytical Balance* (6 min, c, s);
YF-209 *Single Pan Analytical Balance* (6 min, c, s);
YF-210 *Techniques of Operating Single Pan Balances* (20 min, c, s);
YF-211 *How to Make a Washbottle* (10 min, b/w and s);
YF-212 *Fire Safety in the Laboratory* (2 min, b/w and s);
YF-205 *Handling Liquids* (4 min, c, s);
YF-206 *Handling Gases* (3 min, c, s);
YF-204 *Handling Solids* (3 min, c, s);
YF-235 *Patterns of Scientific Investigation* (22 min, c, s);
YF-238 *Phase Demonstration* (1 min, c, s) Three phases of naphthalene;
CORF *Laws of Conservation of Energy and Matter* (16 mm, 8 min, c, b/w, s);
EBF *Heat—Its Nature and Transfer* (16 mm, 11 min, b/w, s);
CORF *The Nature of Energy* (16 mm, 11 min, c, b/w, s);
CORF *The Nature of Heat* (16 mm, 11 min, c, b/w, s);
CORF *Measurement in Physical Science* (16 mm, 13+ min, b/w, c, s.

Loops:

E Ealing Cartridged Super-8: Chemistry Laboratory Techniques (app. 4 min, c);
E 84-0066/1 *Weighing with the Triple Beam Balance;*
E 84-0074/1 *Filtration;*
E 84-0082/1 *Use of the Buret;*
E 84-0090/1 *Use of Pipet;*
E 84-0108/1 *Working Glass;*
E 84-0025/1 *The Bunsen Burner;*
E 84-0058/1 *Weighing with the Triple Beam Balance;*
E 84-0017/1 *Handling Solids and Liquids;*
E 84-0041/1 *Heating Liquids;*
E 84-0033/1 *Heating Solids;*
E 80-3254/1 *Conservation of Mass;*
E 80-2961/1 *Thermal Expansion of Solids;*
E 80-3304/1 *Thermal Expansion of Liquids;*
E 80-3437/1 *Energy Conversion;*
E 80-3445/1 *Conservation of Energy* (Pendulums and pile drivers);
E 80-2769/1 *Conservation of Energy* (Potential to kinetic);
ICF 13130 *Heat is Produced in Different Ways* (8 mm, s-8, 4 min, c);
AFL S-5 *Life of Albert Einstein* (8 mm, s-8, 4 min, b/w).

Filmstrips:

DG 588 *Scientific Measurement: Molecules to Stars;*
DG 565 *Harnessing Solar Energy;*
DG 599 *Power Sources of the Future;*
DG 547 *How We Measure Heat;*
DG 619 *Heat and Temperature—Molecular Energy;*
DG 678 *Chemistry and the Nobel Prize;*
DG 682 *Laboratory Skills in Chemistry.*

Transparencies:

K & E 6120-31 The Conservation of Mass-Energy (1 + 5);
K & E 6120-63 Density (1 + 5);
K & E 6105-9 Units of Measurement (1 + 2);
K & E 6105-10 Relationship between Gram and the Cubic Centimeter (1 + 3);
K & E 6105-11 Use of Powers of 10 (1 + 2);
K & E 6105-14 Weight versus Mass (1 + 3);
K & E 6105-33 Temperature Scales (1 + 2);
K & E 6105-34 Heat Units (1 + 2);
McW The Law of Conservation of Mass;
HRW CH-7101 Equal volume of unlike materials need not have equal mass;
HRW CH-7102 Energy transformation;
HRW CH-7103 One inch equals 2.54 centimeters;
HRW CH-7104 Comparison of English and Metric Systems;
CT 600-37 Temperature Scales;
CT 600-38, 39, 40, Units of measurement (Weight, Length, Volume);
CT 600-49 Physical States of Matter;
CT 605-8 Cancellation of Units;
CT 605-24 Heat and Temperature.

Slides:

W Temperature and Heat; Filtration
W Law of Conservation of Mass; Purification of Water
W Laboratory Skills: How to Heat the Test Tube, Dilute Acids, Use the Balance, Evaporate, Use the Cork Borer, Use the Crucible
W Filtration
W Separation and Extraction
Multimedia programs for individualized instruction: Communications Skills Corp., 1220 Post Road, Fairfield, CT 06430 (filmstrip or slides, audio cassette, student exercise booklets).
CS 801 Converting Numbers to Exponential Notation;
CS 802 Arithmetic Operations with Exponential Numbers;
CS 803 Determining Significant Figures
CS 804 Using Simple Logarithms
CS 805 Solving Simple Equations
CS 806 Collecting and Plotting Data

Chemistry

Vol. 45 (1972) "Can Matter Be Converted to Energy?" Oct.

Vol. 45 (1972))"International System of Units (SI)," Oct.
Vol. 44 (1971) "Science Alone Is Not Enough," Nov.
Vol. 44 (1971) "What Are Scientists Like?" Sept.
Vol. 43 (1970) "Science How It Progresses," May
Vol. 41 (1968) "Nobel Prize Winners," Jan., Dec.
Reprint 111. "International System of Units (SI)," Martin A. Paul, Oct. 1972

Journal of Chemical Education
Vol. 44, (1967) "The Role of Chance in Chemical Investigation," May, p. 299

Science and Children (NTSA), "Heat and Temperature," Nov. 1968

The Science Teacher, "On the Problem of Scientific Discovery, Oct. 1966.

See Scientific American Offprints catalog for articles appropriate to this chapter.

TEACHER NOTES CHAPTER 2

Chapter Notes: Activities and Demonstrations: Classification. Two different approaches may be employed. One is the textbook approach, by which matter is subdivided into categories of materials and further analytic exposition and demonstration are carried on as indicated by Fig. 2-1. Develop Fig. 2-1 on the chalk board. Utilize the equivalent HRW CH-7105 transparency as needed. Develop the figure continually as sections after 2.2 are demonstrated and discussed. The development proceeds either downward to the left hand corner (mixtures) or upward and sideways from the right hand corner (elements). Later, the kinds of elements (metals, nonmetals, metalloids and noble gases) may be placed under the category of elements. A different approach may be termed the synthetic one, in which the reverse procedure is employed. Elements which the student meets in his everyday experiences are identified, and concepts pertaining to the more complex classes of matter are developed. In either approach, extensive use of exhibit materials should be made. For example, after examination of granite with a hand lens, separate large specimens of mica, feldspar, and quartz should be exhibited. Later, sample of the separate elements comprising these three materials should be shown. Schedule Experiment 3, Elements, Mixtures, Compounds. Appropriate portions of Exercise 2. *Exercises and Experiments in Chemistry* may be assigned corresponding questions from the end of the chapter.

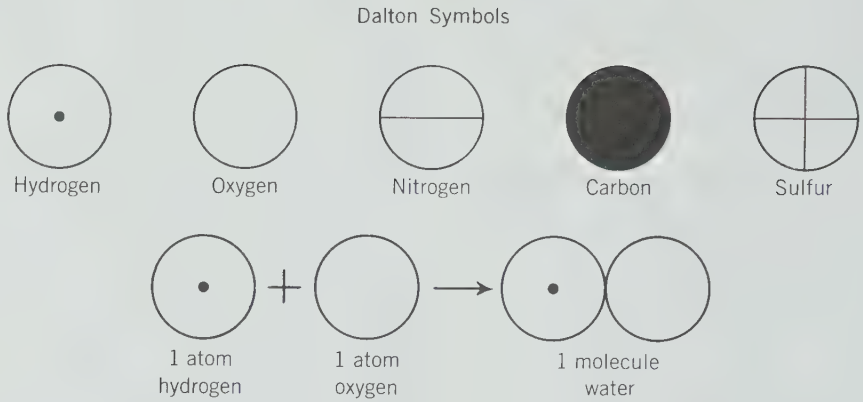
Section 2.2. **Exhibit:** granite, quartz, feldspar, mica. **Demonstrations:** Prepare a sugar solution. Prepare a salt solution. To one half of this solution add water. To the other half add more salt. Students taste solutions with warning about tasting unknown substance. Use distillation and evaporation. Use hand lens to examine resultant salt crystals. **Film:** YF-239, *Solution, Evaporation, and Crystallization.* **Demonstration:** Burn splint in O_2 and in air. Test with limewater solution. Expose a watch glass with limewater solution to air and note white crust deposit with a lapse of time. **Film:** YF-217, *Distillation from Liquid Air.* **Demonstration:** Zn + HCl; Cu + HCl; brass filings + HCl. **Demonstration:** Thermal decomposition of sugar. **Films:** YF-219, *Electrolysis of Water;* YF-216, *Decomposition of Mercuric Oxide.* **Demonstrations:** Expose photographic film. Flash photoflash bulb. **Transparency:** HRW-7150 or C T 605-4, *Classification of Substances* to complement chalk board development of Fig. 2-1.

Section 2.4. Elements (known, synthetic). Exhibit samples of common elements and the Periodic Table Chart. Use the PP transparency T 35 Periodic Table (overlay include Mendeleyev elements and transuranium elements), or K & E 6120-50 Periodic Table of Elements, HM 2-10521 The Periodic Table, C T 600-27 Periodic Table. Student reports may be based on the following references from *Chemistry*, Vol. 43 (1970): "From Mendeleyev to Mendelevium and Beyond," Jan.: "Predicted Properties of Elements 113 and 114," Nov.; "Discovery of Element 105," June; Vol. 42 (1969): "Element 104" May; "A New Element, Eka-platinum?" Feb.; "Henry Cavendish, Discoverer of Hydrogen," Nov.; Vol. 41 (1968): "Origin of Elements in Stars," Apr. Vol. 30 (1957): "Newest Element Christened Nobelium," Sept. Also see *Journal of Chemical Education*, Vol. 36 (1959), "Symposium. The New Elements," Jan. See references to *Scientific American, The Science Teacher, Handbook of Chemistry and Physics, Discovery of the Elements* at the end of this chapter. Also consult the Element profile heading in the index in the Dec. issues of *Chemistry.* Audio-visuals available are: AEC *The Alchemist's Dream*, AEC *The Atom in Physical Science*, FS 376 *Making Elements*, FS 374 *Modern Alchemy*, CHEM 4178 *Transuranium Elements:* filmstrip DG 645 *Evolution of the Elements.*

Section 2.5. Exhibit samples of the most abundant elements, Table 2.1, Table 2.2, and some representative minerals. Discuss Fig. 2-4: regions and dimensions of the interior of the earth.

Section 2.6. Exhibit representative samples of metals, nonmetals, including noble gases (tubes), and metalloids. Such samples may include sheet metals: Cu, Zn, Pb, Al (including foil) to show malleability. Hammer a Pb sheet to show malleability of Pb. Exhibit wire of various metals to show ductility; aluminum pot or pan to show heat conductivity, coated Cu wire to show electrical conductivity. You may wish to use more elaborate demonstrations of heat conductivity (Welch 1653 Conductometer, six-rod form) or some form of lamp, electrodes, dry cell electric setup. Powder rhombic (roll) sulfur in a mortar. Exhibit chunks of antimony and crush in a mortar. Audio-visuals are the film YF-247, *Metals and Nonmetals*, and the K & E transparency set 6120-51, Four Kinds of Elements.

Section 2.7. Use the Dalton Symbols (below).



A page of Dalton symbols occurs on page II in the front of *Foundations of Chemistry*. Alchemical symbols appear in the book *Chemistry*, by Choppin, G. R. et. al. Silver Burdett Corp., 1970, Chapter 1. Contrast the aim (secrecy) of earlier types of symbols and the resulting difficulties in communication with the aim of modern symbols which constitute a sort of international language. See the overprint in Teacher's Edition for activities, audio-visual aids, and references. Also refer to *Journal of Chemical Education*, The "Derivation of the Names of the Elements," Vol. 10, 1933 (pp. 227-234), Vol. 15, 1938 (pp. 123-126): and to *Handbook of Physics and Chemistry*. Another useful transparency is K & E 6105-82, Atomic Symbols. **References:** *Chemistry* Vol. 41(1968) "Elements and the Derivation of Their Names and Symbols," Feb.: Vol. 40(1967) "How and Why of Chemical Symbols," Feb. **Transparencies:** HRW CH-7116 (Fig. 5-3), "Sublevel Blocks of Periodic Table"; PP T35, Periodic Table.

Sections 2.9 and 2.10. Refer to student results in Experiments 3, "Elements, Compounds, Mixtures" if previously performed. Exhibit samples of common mixtures and compounds. Make up mixtures of varying proportions, i.e., salt and pepper, salt and sand, etc. Summarize differences in chalk board replica of Table 2.4 as they become evident. **Demonstration:** Using the Periodic Table have students find the atomic weights of Fe and S and solve the problem: Why is a 7:4 ratio used in forming the compound? On this basis have students attempt to write the formula, since the 1:1 atom ratio will then be clear. Additionally, a qualitative approach to the law of definite proportions may be used. Try three different mixtures of Zn and S: 1 g of Zn to 1 g of S, 2 g of Zn to 1 g of S, 3 g of Zn to 1 g of S, igniting one after another under the hood with the flame of the burner held at arm's length. Note which of the above ratios burns the best and has the least free Zn or S in the product. **Films:** YF-240, *Properties of Mixtures and Compounds;* CORF, *Elements, Compounds, Mixtures;* PSSC 0111, *Elements, Compounds, Mixtures.* **Film:** PSSC 0110, *Definite and Multiple Proportions.* **Loop:** E 84-0165/1, "Definite Proportions: Electrolysis of Water."

Section 2.11. **Exhibit:** Mixtures and compounds in text. **Demonstration (optional):** Na + H_2O, Na burning in Cl_2 (Alyea, p. 201).

Sections 2.12 and 2.13. **Demonstrations:** Perform text examples. In addition, powder sulfur and antimony (mortar and pestle); bend glass. Heat paradichlorobenzene and have students recognize phase changes as sublimed solid solidifies on upper part of tube. Pour the liquid on a cold surface (table top). Have students cite types of energy change (exothermic, endothermic) involved in phase changes. Dissolve solid NH_4Cl or KNO_3 and pellets of NaOH in H_2O with inserted thermometers or have students feel outside of test tubes. For chemical changes refer to previous demonstrations or experiments: FeS formation, burning H_2, synthesis of water, electrolysis of water, decomposition of sugar. Assign selected questions from text and Review Exercise, Unit 1, Chemistry in A Modern World from *Exercises and Experiments in Chemistry* as the basis for a review lesson. Schedule Test 1 from either *Tests* or *Alternate Tests in Chemistry.*

Section 2.14. **Demonstrations:** *Endothermic:* go-stop-go electrolysis of H_2O; thermal decomposition of sugar; go-stop-go decomposition of HgO with splint tests. Exothermic: burn H_2, acetylene, match-paper, burner, S and P in deflagrating spoon, ammonium dichromate.

Have students report on results of Experiment 5, Oxygen: Catalysis. **Laboratory Experiment 5:** Oxygen, Catalysis; Weaver, *Scientific Experiments in Chemistry*, Experiment 4, Heating Metallic Oxides, Experiment 5, A Study of Catalysis.

Section 2.15. **Demonstrations:** ice cubes melting in breaker. Review boiling and evaporation of water. Develop chalk board replica of Fig. 2-9 and later use HRW transparency, Ch-7106 in summary. Demonstrate evaporation of H_2O from moistened cotton wad attached to thermometer bulb to show how higher entropy overcomes unfavorable higher energy. Refer to production of ice cubes in refrigerator where the energy-change factor overcomes the entropy-change factor. Open a bottle of soda water and discuss entropy change. To illustrate disorder-entropy, shuffle a stacked deck of cards, reverse the shuffle. Stir layered colored marbles in a beaker or battery jar. Stir in reverse direction. **Film:** ROB, *Entropy;* ROB, *The Second Law.* **Laboratory Experiment 6:** Energy and Entropy—Phase Change.

Films:

MLA Modern Learning Aids
- 0111 PSSC *Elements, Compounds and Mixtures* (16 mm, 33 min, c, s);
- 0110 PSSC *Definite and Multiple Proportions* (16 mm, 30 min, b/w, s);
- 4178 CHEM *Transuranium Elements* (16 mm, 23 min, c, s);

AEC *The Alchemist's Dream* (free loan, 16 mm, 29 min, b/w, s);

AEC *The Atom in Physical Science* (free loan, 16 mm, 26 min, b/w, s);

INDU FS 376 *Making Elements* (16 mm, 29 min, b/w, s);

INDU FS 374 *Modern Alchemy* (16 mm, 29 min, b/w, s);

AIM Association Instructional Material, c/o Association Films, Inc. Yale Films (16 mm, generally 4 min, c, s).

YF-214 *Catalytic Decomposition of Hydrogen Peroxide;*

YF-240 *Properties of Mixtures and Compounds* (Fe + S → FeS);

YF-215 *Catalytic Decomposition of Potassium Chlorate;*

YF-216 *Decomposition of Mercuric Oxide;*

YF-217 *Distillation from Liquid Air;*

YF-219 *Electrolysis of Water;*

YF-239 *Solution, Evaporation, and Crystallization;*

YF-247 *Metals and Nonmetals;*

FAC Film Associates of California
- *Combustion, An Introduction to Chemical Changes* (16 mm, 16 min, c, s);

CORF Coronet Films
- *Elements, Compounds, Mixtures* (16 mm, 30 min, b/w, c, s);
- *Physical and Chemical Change* (16 mm, 28 min, b/w, c, s);

EBF 1746 *Explaining Matter: Chemical Change* (16 mm, 11 min, c, s);

SUTH Sutherland Educational Films, Inc., 8425 West Third St., Los Angeles, CA 90048

SUTH *Combustion* (16 mm, 14^+ min, c, s);

ROB *Entropy* (16 mm, 30 min, b/w, s);

ROB *The Second Law* (16 mm, 30 min, b/w, s).

Filmstrips:

DG 645 Evolution of the Elements (40 f, c);

DG 556 Oxidation (40 f, c);

DG 609 Catalysts at Work (40 f, c).

Transparencies:

PP Project Physics Overhead Projection Transparencies, Holt, Rinehart and Winston
- T 35 Periodic Table (1 + 7);

K & E Keuffel & Esser Co.
- K & E 6120-50 Periodic Table of Elements (1 + 2);
- K & E 6120-51 Four Kinds of Elements (1 + 5);
- K & E 6120-37 Rate of Reactions—Effects of Catalysts (1 + 9);
- K & E 6105-82 Atomic Symbols (1 + 4);

HM Houghton Mifflin
- 2-10521 The Periodic Table;

C Cambosco Scientific Co., Inc.
- T 605-4 Classification of Substances;
- T 600-27 Periodic Table;
- T 600-51 Energy and Chemical Reactions;
- T 600-52 Energy of Activation: Exo- and Endothermic and Catalyzed Chemical Reactions;

HRW CH-7105 Classes of Matter (Fig. 2-1);

HRW CH-7106 Changes of Phase (Fig. 2-9);

HRW CH-7116 Sublevel Blocks of Periodic Table (Fig. 5-3);

Slides Western Publishing Co., Inc.
- W Mixture and Compound;
- W Decomposition and Synthesis of Water.

Loops:

MH 669026 *Burning a Candle—Combustion I* (8 mm, 1^+ min, c, s);

MH 669027 *Combustion II—Burning Phosphorus* (8 mm, 3^- min, c, s);

ICF 13150 *Conditions Necessary for Combustion* (8 mm, s-8, 4-min, c, s);

E 84-0165/1 *Definite Proportions: Electrolysis of Water* (s-8, 3 min, c, s);

E 84-0173/1 *Combining Volumes: Synthesis of Water* (s-8, 3 min, c, s).

TEACHER NOTES CHAPTER 3

Chapter Note: **Activities and Demonstrations:** A very good source for student reports for Chapters 3 and 4 is the *Project Physics Text*, Chapters 17–20.

Section 3.1. **Demonstrations:** The particulate, nature of matter: Sugar and gas as described in text. Drop a crystal of $KMnO_4$ into a large cylinder of water to show diffusion. Other suitable demonstrations are given

in Alyea, *Tested Demonstrations*, p.63, 5-10-s, p.152, p. 180.

Section 3.2. **Review:** Secs. 2.7, 2.8, 2.10. **Student reports or teacher discussion of:** History of atomic theory through Dalton. Reference—*The Project Physics Text, Unit 5*, pp. 1-15. Present Dalton's postulates from this source. Discuss their use in explaining the laws of conservation of mass and definite composition as outgrowths of experiments about the ways and proportions in which substances react with one another. Help the class "discover" significant differences between Dalton's ideas and the five text statements about atoms and their properties. This leads to questions as to why and how Dalton's ideas were revised, i.e., statements 3, 4 of series 1-5, Section 3.2. **Demonstrations:** This series of demonstrations provides background for an understanding of details of atomic structure before considering Rutherford's scattering experiment.

(a) Charge ebonite and glass or plastic rods by friction. Mount rods in metal saddles and show attraction of unlike charges and repulsion of like charges.

(b) Touch torn paper strips with the charged rubber rod.

(c) Use a suspended pith ball to show attraction, contact, and final repulsion.

(d) Place a charged rubber rod in a box of sawdust or puffed rice. Remove the rod with attached particles and note subsequent repulsion and ejection from the rod.

(e) Charge an electroscope (Graun) by contact and induction. With the aid of a chalk board diagram, have students attempt explanations in terms of the two kinds of electric charges and the mobility of negative charges and/or particles.

(f) Show charged balloons sticking to a wall or a student's back.

Section 3.3. **Slides:** If possible, prepare your own slides of transparencies of Table 3.1 and of the statements 1-5 about atoms and their properties. **Demonstrations:** Show that atoms are not indivisible particles by using a radioactive source and Geiger Counter. Describe Becquerel's discovery. (See: Shamos, *Great Experiments in Physics*, Chap. 15.) To show the complex nature of radioactivity, use a chalk board diagram of Fig. 30-5 or the equivalent transparencies HRW CH-7196, "The effects of a magnet on different types of radiation," PP T-40, "Separation of α, β, γ rays," K & E, "81 Natural Radioactivity." Use charged plates so that students can explain the amount of deflection as well as the charges on the particles.

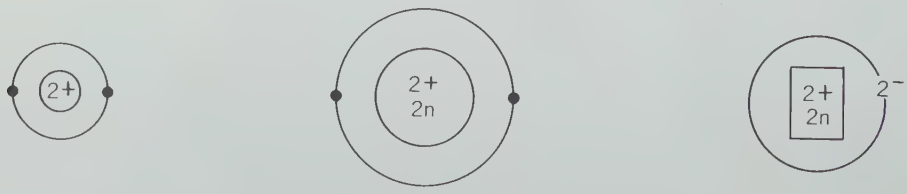
Section 3.3. Rutherford's scattering experiment can be described with the aid of a chalk board diagram (Fig. 3-2). Students should understand that the fold foil is composed of atoms. They should explain the relative numbers of undeflected and deflected particles in terms of the nature of the atom, dimensions of the atom and the nucleus, and the structure of the nuclear atom. The results of the experiment may be interpreted to indicate the type of charge of the nucleus and electrons, the mass of the nucleus, and, later on, the nuclear components.

Section 3.4. Refer students to Section 4.3 to learn how spectra of atoms provide data related to electron arrangement.

Sections 3.6-3.10. **Chalk Board Diagrams:** Types of chalk board models which may be used are represented by those of helium-4 below. Have students construct similar models for helium-3. Develop equivalent diagrams of type with circles for areas of electrons and the central nucleus for hydrogen and helium and for atoms of the second and third series. Develop tables such as in Q.26 of the third series as well as the models. Use the tables and the models to develop concepts and definitions of atomic number, isotope, nuclide, and mass number.

Table 3.2. **Periodic Chart and Transparencies:** Periodic Table; HRW CH-7116, "Sublevel blocks of Periodic Table"; K & E 30-6150 79, "Periodic Chart of the Elements"; HM 2-10521, "The Periodic Table"; PP T39, "Energy Levels—Bohr Theory"; HM 2-10525, "Electron Energy Levels." Use the above to locate atomic numbers, atomic weights, and electron energy levels on the chart and/or transparencies.

Section 3.11. **Definitions:** Emphasize the definition of atomic mass and the basis of the scheme, i.e., carbon-12 = exactly 12. Use Table 3.2 as an aid to definition of mass number as the integer closest to the atomic mass and the total number of protons and neutrons in the atom nucleus. Mention the atomic masses of electron, proton, neutron and relate to the placement of the mass number in their representative symbols. **Demonstration:** Exhibit a container of one dozen pencils and a box containing one gross to aid concept



of *relative* scale. Use a balance to measure the mass of a nickel and then of a quarter. Have students calculate the relative mass of quarter-nickel. Measure the mass of X number of nickels and X number of quarters in small plastic bags. Then have students count and find that X is the same for both coins.

Section 3.12. **Define and illustrate:** Avogardro's number, mole: Mention such units as a dozen used by a grocer (egg carton) and ream used in purchasing paper to illustrate that different types of activities use different unit quantities. Mention mole as the chemist's unit indicating that it has an experimental basis. Use introductory frames of the DG 643 filmstrip, "The Mole Concept." Use simple arithmetic questions: How many atoms are there in **(a)** 2 moles, **(b)** 0.5 mole of carbon-12? How much of a mole of carbon-12 atoms has a mass of 6 grams? Refer to Table 3.3. Have students use the atomic weights in the Periodic Table to obtain the mass number of the most abundant isotope.

Section 3.13. **Define and illustrate:** gram-atomic weight and atomic weight. Refer to Periodic Table chart or transparency. Duplicate the text examples for hydrogen and carbon. Refer students to Tables 4 and 5 in the Appendix. Solve the Sample Problem. Exhibit or weigh out one mole (gram-atomic weight) and then fractional or multiple mole(s) of Zn (powder), S (powder), Hg. Use questions 20, 21, 27, 28, 30, 31, 32. Some teachers may wish to extend the role of prediction based on theory to the problem of the Law of Multiple Proportions predicted by Dalton and then supported by fact and experiment (See *Teacher Resource Book for Project Physics, Unit 5*, p. 15). Teachers may also wish to enrich or complement the textbook treatment in this chapter by using films, the HRW and other transparencies, demonstration experiments, and references for student reports which are listed in the succeeding pages of this section.

Student report references: Kieffer, W. F., *The Mole Concept in Chemistry*, pp. 1–10; Young, J. A., Malik, J. G. "Chemical Queries," *Journal of Chemical Education*, 48, 271(1971).

Films:

The Yale Chemistry Films
- YF-288 *Millikan's Oil Drop Experiment* (6 min, b/w, s);
- YF-254 *Static Electricity, Induction* (9 min, s and c);
- YF-226 *Crooke's Tubes* (8 min, c, s.);
- YF-256 *Principles of Mass Spectrometer* (11 min, c, s.);
- YF-245 *E/M Demonstration* (4 min, s and c).

PSSC Films
- 0416 *Rutherford Atom* (40 min, s);
- 0413 *Mass of the Electron* (18 min, s);
- 0404 *Millikan Experiment* (30 min s);
- 0412 *Electrons in a Uniform Magnetic Field* (10 min, s);
- 0110 *Definite and Multiple Proportions* (30 min, s);
- EBF 1947 *Atomic Energy* (11 min, s, b/w);
- Eye Gate House, Inc., *The Atom* (11 min, s, b/w) 2716 Forty-First Ave., Long Island City, N.Y.;
- DG 616 Atomic Structure and Chemistry (40 f, c);
- DG 643 The Mole Concept (40 f, c).

Transparencies:
- C T 600-5 Atomic Structure and Elements
- C T 600-14 Energy Levels of Electrons
- C T 600-27 Periodic Table
- C T 600-46 Mass Spectrograph—Schematic
- C T 600-47 Mass Spectrograph—Animation
- C T 600-53 Cathode Rays
- C T 600-54 Rutherford's Scattering Experiment
- PP-T-35 Periodic Table
- PP-T-38 Alpha Scattering
- PP-T-39 Energy Levels—Bohr Theory
- PP-T-40 Separation of α, β, γ, rays
- PP-T-45 Mass Spectrometer

Transparencies: See more transparencies on p.55.

K & E (with overlays)
- 30-6150 79 Periodic Chart of the Elements
- 30-6150 78A Model of the Atom
- 30-6150 80 Rutherford's Scattering Experiment
- 30-6150 81 Natural Radioactivity
- 30-6150 82 Atomic Symbols
- 30-6150 85 Mass Spectroscope
- 30-6120 75 Rutherford's Scattering Experiment
- 30-6120 48 Law of Multiple Proportions
- 30-6120 49 Periodic Table and Atomic Structure
- 30-6120 50 Periodic Table of Elements
- 30-6120 51 Four Kinds of Elements

W Color Slides: (2 x 2)
- Q-5 Fluorescence by Cathode Rays
- Q-9, 10 Cathode Rays and Electrical Field
- Q-11, 12 Cathode Rays and Magnetic Field
- Q-13, 14, 15, 16 Deflection of Cathode Rays

HM
- 2-10520 The Nuclear Atom
- 2-10521 The Periodic Table
- 2-10525 Electron Energy Levels

HRW
- CH-7107 Isotopes of hydrogen (Fig. 3-4)
- CH-7108 Helium-4 atom (Fig. 3-5)
- CH-7109 Lithium-7 atom (Fig. 3-6)
- CH-7116 Sublevel blocks of Periodic Table

TEACHER NOTES CHAPTER 4

Chapter Note: Student mastery of the details of the arrangement of electrons in atoms requires that they play an active role, particularly in the methodical, element-by-element evolutions of the various schematics on the chalk board and in-the-seat. Once basic concepts such as electron-pair, octet stability, Hund's rule, and the Pauli exclusion principle have been developed, the students should be able to apply the rules to electron arrangements. Students should use an element's atomic

number and their knowledge of electron configuration from Chapter 3. Students should realize that the Aufbau process of building electron structures of atoms as atomic numbers increase, is a *mental process.* (See the Editor's note on page 66.) Models and schematics of the electron arrangement of atoms in the ground state are the result of this process. Students may not fully understand that they are dealing with the wave-mechanical model of the atom, but any demonstrations or activities that afford them some confidence in the model enhances the eventual usefulness of the model. Some teachers may make use of the ionization data in Toon, Chapter 5 in order to develop or to confirm electron arrangements particularly in the orbitals.

	ev	kcal/mole	
n = 1	-13.6	-313.6	Teachers may wish to use
n = 2	- 3.40	- 78.4	these energy values which
n = 3	- 1.51	- 34.8	relate to the transitions in
n = 4	- 0.85	- 19.6	Fig. 4-4 or transparencies:
n = 5	- 0.54	- 12.5	HRW CH-7110, PP T39 set.
n = 6	- 0.4	- 9.7	

Section 4.1. **Redemonstrate or review:** Electrostatic attraction and repulsion. Frame questions as to how this behavior does and does not satisfactorily explain the structure of atoms and their failure to collapse. Review electron configurations by lettered shells as in Question 26, Chapter 30. Use models as needed.

Section 4.2. **Activity:** Refer interested students to Experiment 43, "Photoelectric Effect" in *Project Physics Handbook.* The demonstration shows that photons of a particular energy produce the emission (loss) of electrons.

Section 4.3. **Demonstrations: (a)** continuous spectrum using a slide projector, a mounted slit, and either a prism or diffraction grating, **(b)** representative flame tests, **(c)** activated gas tubes (H, Ne, Ar); (b) and (c) represent radiation from excited atoms. **References:** Williams, *Modern Physics,* Secs. 12.1-12.8; Castka, *Chemistry Problems,* Chap. 11, Part II, Sec. A, "Spectra and Waves"; Shamos, *Great Experiments in Physics,* "The Quantum Hypothesis," "Max Planck; Toon, *Foundations of Chemistry,* Chap. 6. **Demonstrations:** Line spectra: Students examine the light from a mounted neon tube through a slit and diffraction grating and establish that the emitted radiation consists of definite lines and energy level transitions. Base the development of textbook Secs. 4.2 and 4.3 on the suggested demonstrations in order to explain the wave nature of light, relationships between wavelength and frequency, a photon as a quantum (unit) of electromagnetic radiation energy, and the wave-particle duality of light.

Section 4.5. **Develop:** The concept of orbitals in terms of probability, relating the concept to the Heisenberg Principle. Describe the electron cloud and illustrate it by a chalk board diagram. Define quantum numbers in terms of what each indicates. Use the Pauli exclusion principle to explain why no electron may have the same four quantum numbers. You may wish to employ an analog which compares the electron's quantum number "address" with the details of a student's address such as state, city, street, and street number. Refer to zip code number as a kind of U.S. quantum number pinpointing a location in a certain geographical area. Students should realize that the letters *d, s, p,* and *f* are subdesignations for the orbital quantum numbers. **Exhibit:** models of *s* and *p* orbitals. One source of models is set KC9001, Klinger Scientific Co. **Develop:** A chalk board chart for Table 4.1. Two additional references are: Krupsaw, M., *Journal of Chemical Education,* 49, 433, (1972), "Electron Configuration Diagram"; Grotz, L. C., Gauerke, E., *Chemistry,* 45(1972), "Orbital Memory Devices," May.

Table 4.2. **Activity:** Develop the variety of electron notations in Table 4.2 through student activity. Use a mimeographed worksheet or a suitably arranged student notebook. Discuss one element at a time. Write the results on the chalk board at the same time the students are filling in their notebooks. The information given is the symbol and atomic number of the element. Deal with the first twenty elements to get the pattern of filling in orbitals. (Some teachers go as far as $_{30}$Zn in the tabulation. This introduction increases student understanding through student performances related to part of Table 4.3 and Section 4.7). Work with the principle quantum numbers and electron configuration notation. The generalized formula $n\ell^x$ is developed from the electron notations for H and He. The letter n represents the principal quantum number, ℓ represents the letter sub-designations (*s, p, d, f*) for orbital quantum number, and x represents the electron population of each orbital. If you just use the textbook model, x represents the electron population of the sublevel. If you choose to use the more detailed model which is more directly related to the Orbital notation and Electron-dot notation, the equivalent for $_7$N is $1s^2\ 2s^2\ 2p_x^1\ 2p_y^1\ 2p_z^1$ instead of $1s^2\ 2s^2\ 2p^2$. This activity may take one complete period. Some teachers go as far as $_{30}$Zn in this tabulation, introducing the transition elements (display Periodic Table) and increasing understanding through student performance related to part of Table 4.3 and Sec. 4.7. Some teachers develop the more detailed Electron-configuration notation after the orbital notation equivalents are developed. Emphasize concepts involving the electron-pair, and the octet rule for the outermost electron shells in developing student understanding and mastery of the various notation schemes. Some teachers use the term *valence electrons* (Section 6.2) for the outer-shell electrons.

The orbital notation portion of the Table is developed next. The first problem arises with $_6$C, which is resolved on the basis of electron repulsion of electrons with parallel spins. Hund's rule of maximum multiplicity may

be developed and stated, if desired, as: additional electrons singly occupy equivalent orbitals of the same sublevel before beginning to pair up. The first pairing of *p* electrons occurs with $_8O$.

Develop the electron-dot notation through atomic number 20. Explain the placement of single electrons and electron-pairs about the chemical symbol in the electron-dot notation as representative of the arrangement of the outer-shell (valence) electrons in the orbital notation. Then have the students fill in the electron-dot notations in the last column of the table. Because of multiple valence of the transition elements electron-dot symbols are used infrequently. Have students write electron-dot symbols for the non-transition elements of the fourth and succeeding series. Have students use the data in Tables 4.3, 4.4, and 4.5 in writing these electron-dot symbols. One scheme which students may use to help them in determining the order of energy levels of subshells is shown on p. 276 of *Chemistry Problems* (Sequence of Entrance of "Last" Electron into sublevels). If the transition elements are included in the classroom table, as it develops, the various configurations are developed showing that though additional electrons are present *after* Ca(4th series) they are in the 3*d* sublevel after 4*s* (Aufbau buildup), as shown in Fig. 4-8. However, in the atom in its ground, the 3*d* electrons are below the 4*s* electrons (Table 4.3) for Sc-Zn. The orbital notations for Cr and Cu may be used to develop the concepts of extra stability of half-filled (Cr) or completely filled (Cu) orbitals (sublevels). The completion of the 4*p* orbitals after completion of the 3*d* orbitals is illustrated by reference to Table 4.3. The arrangement of the 4*f* orbitals after the 6*s* orbitals are filled is followed by the completion of the 5*d* sublevel (Table 4.5).

Demonstration and review: Use Fig. 5-4 and the HRW **Transparency** CH-7117, "Sublevel Blocks of the Periodic Table" to show how elements may be arranged according to their atomic structure (Refer to Tables 4.3, 4.4, 4.5, 4.6). Compare the electron-dot symbols for Groups I, II, III–VIII elements. Review previous treatment of arrangements of *d* and *f* electrons and the schematics for placement. Use questions 18–20. Then select from questions 21–28.

Audio Visual Materials

Films:

4148 CHEM *The Hydrogen Atom—As Viewed by Quantum Mechanics* (standard version) (13 min, s and c);
4149 CHEM *The Hydrogen Atom—As Viewed by Quantum Mechanics* (advanced version) (20 min, s and c);
CORF *Light, Wave and Quantum Theories* (16 mm, c or b/w, s, 13 min);
INDV FS-358 *Why Are Atoms Unpredictable?* (16 mm, b/w, s, 30 min);
INDV FS-357 *What Makes Atoms Stick Together?* (16 mm, b/w, s, 30 min);
0415 PSSC *Electromagnetic Waves* (33 min, s);
0417 PSSC *Photoelectric effect* (28 min, c, and s);
0418 PSSC *Photons* (19 min, s);
0419 PSSC *Interference of Photons* (13 min, s);
0423 PSSC *Matter Waves* (28 min, s);
0404 PSSC *Millikan Experiment* (34 min, s).

Filmstrips:

DG 609 Orbitals—Atomic and Molecular (fs, c, 40 frames), frame 10-24
DG 629 Orbitals—Atomic and Molecular;
EB 919-97 Atomic and Molecular Models.

Transparencies:

T39 Project Physics Energy Levels—Bohr Theory;
K & E 6120-25 Diagram of *p* and *d* orbitals (6 overlays);
K & E 6120-26 Periodic Table and Atomic Orbitals (4 overlays);
C T 600-4 Atomic Orbitals;
C T 600-11 Electromagnetic Spectrum;
C T 600-13 Electron Distribution Wheel;
C T 600-14 Energy Levels of Electrons;
C T 600-15 Energy Difference of Electron Levels Resulting in Spectral Lines;
C T 600-32 Quantum Designations of Electrons;
C T 605-13 Hydrogen Spectrum Lines and Electron Energy Levels (no overlays);
HM 2-10524 Electron Orbitals;
HM 2-10525 Electron Energy Levels.

Slides:

Western Publishing Co., Inc.
Spectra
Color
The Flame Test

TEACHER NOTES CHAPTER 5

Chapter Note: The primary objective of this chapter is to help students develop demonstrable appreciation and skills concerning both the sequence of experimentally discovered facts about atomic structure and the organization of these facts. Correlated with previously developed models of atomic structure (Chapters 3 and 4), such skills can unify and simplify the study of chemistry. This objective also involves awareness that by using accepted generalizations, reasonably accurate predictions can be made.

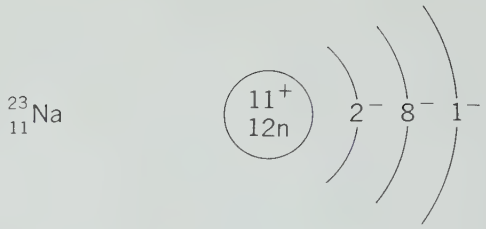
Section 5.1. **Transparency:** PP T35, A, B, C, D, E, F, G, H. **Alternative Exhibit,** wall chart of periodic table with covers over noble gases; Sc, Ga, Ge; series 6 and 7. Discuss history of attempts at classification (Dobereiner's triads, Newland's law of octaves). **Demonstrations:** Similarity in properties **(a)** Li, Na, K on H_2O, **(b)** powdered Sb (cold) into Cl_2, Sb (heated) into Br_2, Sb (heated into hot I_2). Students compute atomic weight of Br as approximately one half of sum of those of Cl and I. Students count off number of elements before property reoccurs: Li to Na to K. Have students explain why Mendeleyev table had periods of 7 and 17

instead of 8 and 18 (remove cover from noble gases or use T35E). Help students discover discrepancies of placement by atomic weights of Ar versus K and Te versus I. Note that Mendeleyev resolved this by using the properties rather than by using the atomic weights which he thought were in error (although many actually were very close to presently accepted values). Have students attempt to explain other gaps, how Mendeleyev handled the difficulties and predicted missing elements and their properties, "Eka-silicon" and the experimentally reported properties in 1886 (atomic weight 72 versus 72.6, color dark gray versus grayish white). **References:** Brescia, pp. 167–172; *Project Physics Text*, Secs. 17.3–17.7, pp. 16–25. **Transparencies:** C T 600-21, Periodic Table; HM2-10521, The Periodic Table.

Section 5.2. **Discussion:** Have students explain how atomic numbers, the basis of the modern Periodic Table, correctly place Ar and K, and Te and I, because atomic numbers are a more fundamental property than atomic weights. **References:** For student report on Moseley, Snyder pp. 69–70, Jaffe, *Moseley and the Number of the Elements.* Emphasize the use of "new" apparatus in discovery of facts which lead to new ideas. Mention the mass spectrograph and use **Transparencies:** C T 600-64, "Mass Spectrograph Schematic"; C T 600-47, "Mass Spectrograph Animated"; PP T45, "Mass Spectrometer"; K & E 30-6150, "Mass Spectroscope." Give examples of "missing" elements according to Moseley's calculations (atomic numbers 43, 61).

Section 5.3. **Review:** Demonstrate similarities of Li, Na, K; Cl, Br, I to show that elements of similar properties occur at certain intervals and that properties of elements go through a pattern of change. Also mention the Noble Gases. Have a student state the periodic law.

Section 5.4. **Exhibits and transparencies:** Display a wall chart of the periodic table or one of the previously mentioned transparencies in conjunction with the text table and have students locate atomic number and weight. Have them also note the electron energy levels and electron population. Period and group are defined and the representative numbers identified and located on the table. Have students identify the main types of elements labeled in the table. It is useful at this point to concentrate on the first 20 elements. Exhibit samples of elements of period 3 (Fig. 5-3) and as many of the other elements as possible. Review the electron configuration of elements 1–20 using either chalk board diagrams of the main shell structures or electron-dot symbols to point out similarities in the outermost shell population among members of the same groups (I–VIII). A mimeographed worksheet consisting of boxed outlines for the first twenty elements also affords space for inclusion of the electron configuration notation. See the example at the top of the next column.



Na· $1s^2\,2s^2\,2p^6\,3s^1$

The given data is the symbol with the atomic number and the mass number of the most abundant isotope. Have students draw models and fill in boxes complemented by chalk board development. Use the HRW transparency, CH-7116 (Fig. 5-4) Sublevel Blocks of the Periodic Table. This is also useful for objectives 12–13. Have students put in the zig-zag lines on their worksheet. **Demonstrations:** select demonstrations from the ones listed below to help provide actual experiences in identifying physical and chemical properties and help maintain student interest in the descriptive details which appear in Sec. 5.4.

(1) *Metallic luster:* Na, Mg, Al, K, Ca.

(2) *Toughness of metals versus brittleness of solid nonmetals:* cut Na, K, Mg ribbon, sheet Al; beat Al or Mg with a hammer, attempt to cut S; hit crystal of rhombic (roll) S and powder in mortar.

(3) *Electric conductivity:* Use 2 dry cells in series with a flashlight bulb and place samples across open terminals. By replacing the bulb with a lecture table ammeter and by using the appropriate range with a suitable resistor in series, it is possible to show decrease in conductivity for elements of period 3. A conductivity tester (Welch 2369A), may also be used with 120-v ac and a 25-w bulb.

(4) *Reaction with O_2 (air):* Have students observe storage of Na in hydrocarbon and P in water, rapid tarnishing of freshly cut Na, burning Mg ribbon, burning Na in deflagrating spoon, Al powder in burner flame, "smoking" of white P, burning of S and red P in deflagrating spoon.

(5) *Reactions with H_2O:* Test Na, Mg, Ca, K, C, S, Si. Use the CHEM Loop "Alkali Metal Reactions with Chlorine and Water."

(6) *Reactions with dil. HCl:* Mg, Al, S, C, Si, P.

(7) *Reactions with dil. NaOH:* Test Mg versus Al, nonmetals.

(8) *Identify He, Ne, Ar:* by using gas tubes with Telsa coil.

(9) *Oxides with H_2O and effect on litmus:* Use CaO, SO_2, or P_4O_{10}. The oxides produced by burning S or red P in a deflagrating spoon in a bottle of air may be used along with solid CaO. **Films:** YF-247, *Metals and Nonmetals;* CORF, *Metals and Nonmetals.* **Activities:** The variation in properties across a period is established as follows: Metals of decreasing activity (e.g. Na versus Mg) → metalloids → nonmetals of increasing reactivity → noble gases (inert for He, Ne, Ar). Correlation of class of element with outer-shell electron population is established by comparing representative worksheet models and with element properties. Metals are recognized as having a few electrons—one, two, and sometimes three (Al)—in outermost shell, while the most active nonmetals have an almost complete octet and usually many electrons—five, six, or seven. The inertness of the first three noble gases is related to the stability of an electron pair (He) and of the octet (Ne and Ar). Show the film CHEM 4112 *Chemical Families,* and use it in connection with group properties and variations in properties. It emphasizes the relation of the most active metals and nonmetals as having, respectively, one more and one fewer outer shell electrons than related noble (inert) gases which have an octet. The film also demonstrates the reaction between xenon and fluorine. On the basis of demonstrations such as the reaction of Ca versus Mg with H_2O, Na on H_2O, and K releasing H_2 from ice, and the relative activity of metals in a family is established as increasing with increasing atomic number. Using the reactions of Sb with Cl_2 and Br_2 to establish that chlorine is more active than bromine and, by analogy, to have students predict that fluorine is the most active. Have students generalize that nonmetallic reactivity decreases with a decreasing in atomic number. The filmstrip DG 647 *The Noble Gases* presents evidence of the inertness of He, Ne, and Ar, and of compound formation by Xe and Kr. Student reports on compounds of the Noble Gases may be based on references listed at the back of the Teacher's Section in the front of the text. Identify and discuss the transition elements of periods 4 and 5 and review the electron configurations involving the 3*d* and 4*d* orbitals. Exhibit some of the more common and recognizable metals: Fe, Cr, Cu, etc. Have students explain why the transition elements are metals, emphasizing the electron populations of the outermost *s* sublevel. The first ten frames of the DG filmstrip 633 *The Transition Elements* may be used for an introduction to transition metals. You may also refer students to Table 7 in the Appendix or to the first few pages of Chapter 25 of this text. The electron configurations of the Lanthanide and Actinide series are established as involving incomplete 4*f* and 5*f* orbitals respectively. Point out the similarity of the two outer shells of the rare earth elements in explaining why these elements have almost identical chemical properties. Discuss the significance of the zig-zag line separating metal from metalloid in the Periodic Table. Identify and define metalloids. Discuss the change in properties of groups IV and V: nonmetal → metalloid → metal. The students probably are most acquainted with the Group IV elements, such as C (nonmetal in some forms), (metalloid), Sn and Pb (metals). Note that for groups I–VII, as the atomic number increases, chemical reactivity as a metal increases while chemical reactivity as a nonmetal decreases. Items 1–9 and 19–23 is Exercise 5 of *Exercises and Experiments in Chemistry* may be used as a summary of Sec. 5.4.

Section 5.5. **Activity:** Have students examine the tabular and graphic data first for size of atoms, then for ionization energy, ionization energy to remove successive electrons, and electron affinity, found in Secs. 5.5 to 5.8. They should be challenged and helped to recognize and state generalizations (stated in the text) on the basis of these data and their relationship to atomic structure and the periodicity of properties. **References:** *Journal of Chemical Education,* 45, 791(1968) "The Correct Size of the Noble Gas Atoms"; Sanderson, *Inorganic Chemistry,* "Chemical Periodicity." **Develop** the generalizations and explanations for 1 and 2 p. 84. **Charts:** Welch 4844, Relative Sizes of Atoms and Ions in the Periodic Table; Toon, *Foundations of Chemistry,* p. 247.

If you compare this table of atomic radii (Fig. 5.5) with previous editions of *Modern Chemistry* or with some of the published charts that are available, you will notice differences between them. These differences occur because measuring atomic radii experimentally presents problems for several reasons, among them: 1. The electron does not exist in an orbit at a fixed distance from the atom's nucleus. 2. It is not possible to isolate and measure experimentally the distance from the nucleus to the outermost electrons of a single atom. The only data available results from the average behavior of a large number of atoms under specified conditions, i.e. a covalent compound. For these reasons, the values reported depend upon the method of measurement. If you feel that an explanation of these differences would help your students, refer to a more advanced text, e.g. Mortimer, *Chemistry, A Conceptual Approach.*

Section 5.6. **Films:** CHEM 4151, *Ionization Energy;* PSSC 0421, *Franck-Hertz Experiment.* **Filmstrip:** DG 648, "Ionization Potential and Electronegativity." The attraction between nucleus and electrons leads to the problem: "How can the electron(s) be totally removed from an atom?" The CHEM film shows students how ionization energy is experimentally determined. On the basis of this film and the HRW **transparencies** CH-7118 (Fig. 5-7) and CH-7118, develop the series of generalizations pertaining to the first ionization energies

in Sec. 5.6. Develop functional concepts related to ionization energy for noble gas, metal, nonmetal, and metalloid. The noble gas data furnish evidence for the stability of the octet, and for He of the electron-pair. The general decrease in ionization energy within a group with increasing atomic number is explained by the increase in radius. The general but irregular increase in ionization energy in a period (Groups I–VIII) is identified as another periodic property. Present the explanations in Sec. 5.6 for the decreases between Groups I and II and Group V and VI as they relate to the electron population of sublevels. Another viewpoint is to consider the ionization energies for Be and N (period 2) and for Mg and P (period 3) as being unexpectedly larger than they would be if there were a regular, instead of an irregular increase in ionization energies. From this viewpoint, the stability of the electron-pair and the extra stability of the half-completed sublevel may be developed. **Review:** Electron configuration notations or orbital notations for the elements. Have students explain how the experimental data presented throughout Sec. 5.6 reinforce the models developed on the basis of wave mechanics. Have students locate an unusually large decrease in ionization energies of successive elements in periods 4, 5, and 6. Have them explain this decrease in terms of the population of the $(n\text{-}1)d$ sublevel of np orbitals.

Section 5.7. Discussion and experiments: Have students explain the differences in the values for removal of successive electrons for Na, Mg, and Al in terms of the electron configuration and orbital designation for the electrons undergoing removal. Have students attempt to answer question 28. Some teachers present the ionization energy experiments and data while they are developing the concepts of energy level, sublevels, and orbitals in Chapter 4. Students can, and do, understand the numbers indicating experimental measurements that are reproducible more easily than the abstractions of the wave mechanics.

Section 5.8. Demonstration: Electron affinity is defined, represented by the equation, and the electron transfer is recognized as an exothermic reaction in contrast with ionization electron transfer which is endothermic. The high electron affinity of the halogens is explained on the basis of completion of the octet in the outermost energy level. You may wish to demonstrate the burning of Na in Cl_2 to set the stage for the next chapter and challenge students to attempt to use the concepts of ionization energy, electron affinity, and octet completion in explaining the formation NaCl. **Activity and programmed instruction:** The ideas developed throughout this chapter may be reinforced by using Exercise 5, The Periodic Law and the programmed instructional references at the end of this chapter.

Films:

CHEM 4151 *Ionization Energy*, (22 min, c and s)
CHEM 4112 *Chemical Families*, (22 min, c and s)
CORF *Metals and Nonmetals* (16 mm, b/w, c, s, 13 min)
PSSC 0421 *Franck-Hertz Experiment* (30 min, s)
YF-244 *Electronegativity* (16 mm, c, s, 4 min)
YF-247 *Metals and Nonmetals* (16 mm, c, s, 9 min)
YF-255 *Melting Point Relations to Periodic Table* (16 mm, c, s, 3 min)

Filmstrips:

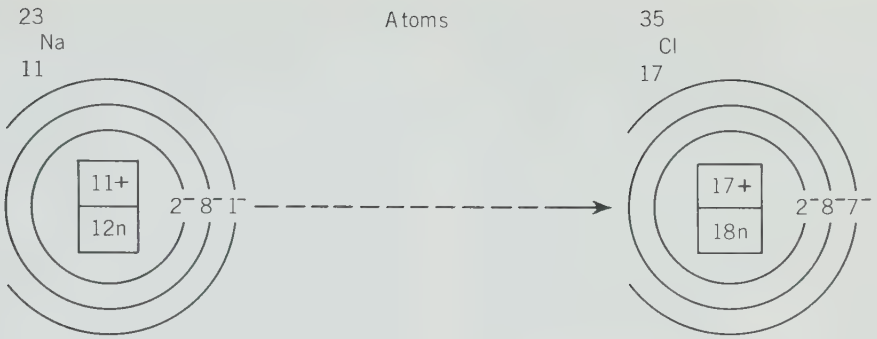
ID-G 647 *The Noble Gases (f*s, 40 frames, c)
ID-G 648 *Ionization Potential and Electronegativity* (fs, 40 frames, c).

Transparencies:

HRW CH-7115 "Periodic Table of Atomic Radii" (Fig. 5-5)
HRW CH-7116 "Sublevel Blocks of the Periodic Table (Fig. 5-4)
HRW CH-7117 "Periodic Table of Ionization Energies (Fig. 5-7)
HRW CH-7118 "Graph showing first ionization energy as a function of atomic number" (Fig. 5-8)
K & E 6120-50 "Periodic Chart of the Elements" (1 + 2 overlays)
K & E 6120-51 "Four Kinds of Elements" (1 + 5 overlays)
K & E 6120-17 "Sizes of Atoms" (1 + 2 overlays)
HM 2-10521 "The Periodic Table"
HM 2-10522 "Trends Within the Periodic Table"
HM 2-10526 "Trends in Ionization Energies"
HM 2-10523 "Some Trends in the Properties of the Third-Row Elements"
C T-600-2 "Atomic and Ionic Diameters"
C T-600-3 "Atomic and Ionic Radii"
C T-600-18 "Ionization Potential"
C T-600-21 "Periodic Table"
PP T35 "Periodic Table" (long form, Elements Known to Mendeleyev, Alkali Metals, Halogens, Noble Gases, Transition, Radioactive and Transuranic Elements.)
W 2 x 2 slides, "The Ionization Tendency of Metals"

TEACHER NOTES CHAPTER 6

Section 6.1. Activity: List the formulas in Sec. 6.1 and have students recognize and propose the problem: "Why are there differences in the number of H (and Cl) atoms per atom of the other elements?" **Laboratory Experiment 6,** "Combining Number," *Scientific Experiments in Chemistry.* Students may be able to suggest a mode of attack based on their knowledge of atomic structure and electron configuration of the atoms. They may recognize that five of the compounds consist of two nonmetals and three compounds consisting of a metal combined with a nonmetal. **Exhibits:** A crystal of halite, crystals of NaCl (hand lens), and both ball-stick (Klinger KC80604) and close-packed (Science Related Materials 7027). Also see Ealing's "Atomic Model Reference Library." Pose the question: "How can the formula NaCl represent the compound and the combined atoms (ions)?" Define empirical formula. Use the orbital configurations for Na, Cl, and NaCl in Sec. 6.3. Use Table 6.1 and the electron-dot formula for NaCl.



Complement these with the atom shell chalk board diagrams with the half-reaction equations, the energies involved, and the labeling of the half-reactions as oxidation and reduction. Use the chalk board first for developing the atom shell model at the top of the page to show how NaCl forms by electron transfer resulting in ions having outer octets. The space underneath the model may be used subsequently to write the equations and identification of the process as oxidation or reduction as energies and processes are discussed later. This model is also useful in counting the protons and electrons in corresponding ions and thus helping students appreciate how ion charge originates.

$Na^0 + \text{energy} \rightarrow Na^+ + e^-$ (Ionization energy) (Oxidation)

$Cl^0 + e^- \rightarrow Cl^- + \text{energy}$ (Electron affinity) (Reduction)

Section 6.4. **Develop:** text questions on the chalk board and the energies involved resulting in the exothermic reaction.

Section 6.5. **Develop:** the three rules for oxidation numbers as they apply to the formation of NaCl. If you used the chalk board atom shell models suggested above, label half-reaction equations. Also define oxidizing agent and reducing agent and label Cl and Na respectively. **Demonstrations:** Using atom shell models, show the formation of MgO and relate to the demonstration burning Mg ribbon, which students readily recognize and label as the process of oxidation. Demonstrate the burning of Mg in Cl_2 (**Caution:** use goggles and class shield) and show the formation of $MgCl_2$ using atom shell models on the chalk board. Have students recognize the similarity for the half reaction: $Mg - 2e^- \rightarrow Mg^{++}$ for both formation of MgO and $MgCl_2$. This permits the framing of the definition of oxidation as electron loss. The gain of electrons by both the O and the Cl is labeled reduction (some teachers also use deoxidation). **Exhibit:** samples or models and place the number, zero, above each of the symbols or formulas. **Develop Rule 3:** by placing the oxidation numbers in the formulas of compounds:

$\overset{0}{Na}\ \overset{0}{Cl}$, $\overset{+1}{Na^+}\ \overset{-1}{Cl^-}$; $\overset{+2}{Mg}\ \overset{-2}{O}$, $\overset{+2}{Mg^{++}}\ \overset{-2}{O^{--}}$. Repeat for $MgCl_2$.

Section 6.6. **Activities:** The preceding activities readily lead to the details in the formula:

$\overset{+2}{Mg^{++}}\ \overset{-1}{Br_2^{--}}$

which appears in the final paragraph of Sec. 6.6. Develop the orbital model representing the electron transfer in the formation of $MgBr_2$. Use Table 6.2, developing details on the chalk board. Refer students to Fig. 6-3. Use the chalk board to develop the electron-dot and ionic formulas for $MgBr_2$. Write the oxidation and reduction equations and label oxidizing and reducing agents. Reemphasize that all formulas for ionic compounds are empirical formulas. You may have students construct atom shell diagrams or electron-dot formulas for $CaCl_2$ and Al_2O_3.

Section 6.7. **Exhibit:** $FeCl_2$, $FeCl_3$, or compounds in which Mn has different oxidation states. Write the formulas for $FeCl_2$ and $FeCl_3$. Have students decide on the oxidation number of Fe in each compound using the oxidation number of the chloride ion (−1) as the basis. Students may recognize that unlike the metals of Groups I and II and Al, the transition metals (HRW transparency CH-7116, "Sublevel Blocks of the Periodic Table") generally have several oxidation states. Your wall chart of the Periodic Table may show the several oxidation states of various elements (Sargent-Welch 4854, Sargent-Welch student chart S-18806). For Fe, you may wish to develop the orbital notation and use the ionization energies per mole: 1st = 182 kcal, 2nd = 375 kcal, 3rd = 708 kcal. The wall chart may also show that nonmetals generally have several oxidation states. **Exhibit:** the Sargent-Welch-chart 4844, the HRW transparency CH-7120 is a replica of this chart. **Development:** Have students use the text two-page table, Periodic Table of Atom and Ion Sizes. Define cation and anion. Develop the generalizations and explanations pertaining to these, including variations within a group and period.

Section 6.9. **Review:** Ionic bonding. Exhibit crystals of other ionic compounds, $MgBr_2$, KCl, NaBr, etc. Discuss the strength of the ionic bond in terms of the

relatively high melting point (a student observer of the melting process describes its melting). Place the halite crystal across the electrodes or insert electrodes into the small crystals of NaCl in a beaker. The failure to conduct in the solid phase also indicates strength of the bond. Crush crystals of NaCl in a mortar to indicate brittleness of ionic crystals. You may wish to demonstrate that water does not conduct but that the solution of the solid NaCl in water does conduct an electric current. Introduce the concept of dissociation and free-moving ions in the liquid phase and in aqueous solution. A difference in the reaction of an ionic and a covalent compound can be shown by using Goldfish, D. M., Journal of Chemical Education, 46, A497 (1969), "Ionic Versus Covalent Bonding."

Section 6.10. **Demonstrations:** Assemble models of HCl, H_2O, NH_3 and use the chalk board for representative orbital configurations and formulas. Have students describe the shapes of the assembled models. Start with the models of the elementary molecules and take (pull) apart to represent bond breaking while the assembly represents bond making. Relate bond breaking to bond energy. In the electron-dot formulas of these three compounds, the bonding electron pair(s) may be placed closer to the central atom than to the hydrogen atom(s). Students can explain this on the basis of the larger nuclear charges of Cl, O, and N. This treatment provides the basis for the later development of bond polarity, dipoles, and electronegativity. It also helps in developing oxidation number rules 4 and 5. **Demonstration:** of bond energies in HCl and HI. A heated glass rod is inserted into HCl(g) with no evident change. This is repeated for a test tube containing HI(g). The heated rod causes decomposition with the production of violet vapors of I_2. A hot iron rod may even cause explosion of the liberated H_2. See Fowles, *Lecture Experiments in Chemistry*, pp. 270-272 for preparation of HI and decomposition by heat.

Section 6.11. **Explain:** Resonance (as in Sec. 6.11 for SO_2) and define the term. You may wish to circle the coordinate covalent bond and indicate that the bonding electron pair is from one atom only, i.e., S. A structural formula (dash bond) which may be used to represent the coordinate covalent bond is:

Section 6.12. **Activities:** Have a student write the electron-dot symbol and the electron-configuration for carbon in the ground state. Indicate that the experimental evidence shows that the bond angles (Fig. 6-6) and bond lengths in the compound CH_4 are equal. Exhibit the ball-stick model, a model of a tetrahedron in clear plastic (Welch 7579), and have students construct a cardboard or soda straw tetrahedron model at home. This approach poses the problem of how carbon can have four equivalent bonds and four equivalent bonding orbitals. **Transparencies:** and loops are used as class progress permits. **Reference:** Refer students to Castka, *Chemistry Problems*, pp. 348-356, particularly pp. 351–352. **Activity:** For H_2O challenge students to write an electron-dot linear structure for H_2O. Draw the *p* orbitals for O on the three axes and show how the overlap of the *s* orbitals of H with the two 3*p* orbitals of oxygen result in a bent (theoretically 90°) shape for H_2O. Relate to HRW **transparency** CH-7124 (Fig. 6-8) and discuss the unshared electron pair repulsion as in Sec. 6.12.

Section 6.13. **Exhibit:** samples of monatomic, diatomic, and more complicated substances and write their formulas (symbols for the noble gases). Refer to Fig. 23-1 for P_4 and Fig. 23-2 for S_8. A model of P_4 may be made from four plastic spheres arranged into a tetrahedron. A model of S_8 may be made from 8 C atoms in the Welch ball-stick models by using only two of the holes in each sphere for connection by sticks. Write various symbols and formulas with coefficients and have students explain what each means.

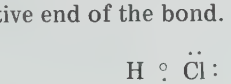
Section 6.14. **Discuss and refer:** Cite numbers representing molecular sizes. Interested students may be referred to Experiment 34, The Size of a Molecule, Lehrman, *Scientific Experiments in Physics*, Holt, Rinehart and Winston, 1968.

Section 6.15. **Activity:** Refer to atom shell models of H_2 and F_2 and write the electron-dot formulas with the bonding electron pair equidistant from each symbol. Ask, "Which nucleus has the greater attraction for the bonding electron pair?" Define pure (nonpolar) covalent bond. Write the electron pair in HCl nearer the Cl. Have students explain that Cl has the greater attraction. Define polar covalent bond. Redefine ionization energy and electron affinity. **References:** for electronegativity scales: Pauling, L., *The Nature of the Chemical Bond*, pp. 63–72; Gould, *Inorganic Reactions and Structure*, pp. 139–141; Sisler, *Electronic Structure and Properties and the Periodic Law*, Table 2.6 Electron Affinities; Toon, *Foundations of Chemistry*, Tables 7.6, 7.7. **Develop generalizations:** for electronegativities from data in Fig. 6-10.

Section 6.16. **Activities:** Use Table 6.6 for calculations of electronegativity difference and identification of bond type. The Welch student chart of Periodic Table has the data for Table 6.6. Use questions 24 and 35. Develop oxidation Rule 6 by referring to the polar bond in HCl. Use electronegativity difference to establish ionic nature of NaH. Pose the problem of bonds in F_2O.

Section 7.17. **Demonstrate:** the test for polarity of liquids. Use two burets, one containing water and the other CCl_4, benzene, or toluene (*Caution*). The thin stream of liquid water from the buret is deflected by a charged rod. It is also deflected by a positively charged rod (glass or plastic). The nonpolar liquid(s) are not deflected. Establish possible polarities of molecules

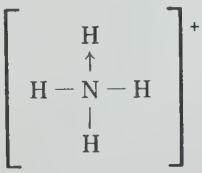
having (a) nonpolar bonds only, (b) polar bonds. Refer to shapes of molecules and the effect of the arrangement (symmetry) of polar bond arrangements as in CH_4 and CCl_4; H_2O and CO_2. Refer back to atom shell models of H_2 and F_2 and write the electron-dot formulas with the bonding electron pair equidistant from each symbol. Ask the question, "Which nucleus has the greater attraction for the electrons or electron pair? Why?" Define pure (nonpolar) covalent bond. Write the electron-dot formula for HCl but place the electron pair nearer the Cl. Have students explain why the Cl has the greater attraction. The electron affinity for H is 0.7. Students may use the electron affinities in the explanation. Refer again to p. 26, Table 2.6 (Electron Affinities) in Sisler's *Electronic Structure Properties and the Periodic Law.* On the basis of the unequal sharing have students define polar covalent bond. One scheme for representing a polar bond consists of an arrow with a plus sign at the tail end, the arrow head representing the negative end of the bond.



The use of this vector type of symbol aids students in understanding how a nonpolar molecule (substance) may have polar bonds. Refer to Toon, *Foundations of Chemistry*, Fig. 8-45, Fig. 8-46 which illustrate the use of such vector representations (BeF_2, H_2O, NH_3). Place the unshared electron pair on top of the N in the spatial structural formula of NH_3 in Sec. 6.17 to help students realize its role in molecular polarity.

Section 6.18. **Demonstrate**: Draw the electron-dot representations for the polyatomic ions in Sec. 6.18. Circle the coordinate covalent bonds in these diagrams. Use the dash and arrow structural equivalents. Locate the coordinate bonds in the polyatomic ion models after drawing them on the board. Circle the coordinate covalent bonds. Use the dash and arrow structural equivalents.

For NH_4^+:



Some teachers use the following demonstrations for an early introduction to acids and bases (protolysis) by employing the coordinate covalent bond as a device for establishing the production of such ions as H_3O^+, OH^-, and NH_4^+ as well as showing that polar compounds may form ions by dissolving in polar water with the solution electrical conductivity.

Section 6.19. **Restate**: the oxidation rules and use questions 25, 37-39.

Films:

MH 611999 *The Structure of Water* (16 mm, 14 min, b/w,s; or 612024 in color);
612007 *Oxidation—Reduction* (16 mm, 9+ min, b/w, s; or 612019 in color);
CORF *Chemical Bond and Atomic Structure* (16 mm, 16 min, b/w, c, s);
YF-224 *Electronegativity* (16 mm, 4 min, c, s);
CHEM (MLA) 4157 *Chemical Bonding* (16 mm, 16 min, c, s; 4857:8, c, s);
CHEM 4057 *Teacher Training Introduction to "Chemical Bonding"* (16 mm, 14 min, b/w, s);
CHEM 4154 *Shapes and Polarities of Molecules* (16 mm, 18 min, c, s); or 4854:8, c, s).

Film Loops:

LGC E-65 *The Properties of An Ionic Compound* (8, 3 min, c);
E-65A *The Structure of An Ionic Crystal* (8, 3 min, c);
E-66 *The Properties of a Covalently Bonded Molecule* (8, 3 min, c);
E-66A *The Structure of the Covalent Molecule* CCl_4 (8, 3 min, c);
LBF 6 *Atomic and Bonding Orbitals* (8, 4 min, c);
MH *Atoms to Molecules* (8 film series) (s-8, 4 min, c);
Q-65-521 *Atomic and Bonding Orbitals* (8 mm, s-8, 4 min, c);
HR *Simple Molecular Orbitals* (s-8, 4 min, c);
E 80-2041/1 *Paramagnetism of Liquid Oxygen.*

Filmstrips:

DG 625 Crystals and Their Properties (40 f, c);
648 Ionization Potential and Electronegativity (40 f, c);
627 Covalent Bonds—Covalent Structures (40 f, c);
629 Orbitals—Atomic and Molecular (40 f, c);
6505 The Ionic Bond (40 f, c);
6506 The Covalent Bond—Molecular Orbitals (40 f, c);
6507 The Coordinate Covalent Bond (40 f, c).

Transparencies:

6120-14 Ionization Reactions in Water—Liquid Solution (2 + 1);
6120-15 Ionization Reactions in Water—Gas (2 + 2);
K & E 6120-16 Hybridization (1 + 5);
6120-17 Sizes of Atoms (1 + 2);
6120-18 Ionic Compound—Lithium Fluoride (1 + 5);
6120-19 Ionic Compound—Magnesium Oxide (1 + 5);
6120-20 Covalent Compounds—Nitrogen Molecule (1 + 3);
6120-21 Covalent Compounds—Ammonia (1 + 4);
6120-22 Covalent Bonding—Ethane, Acetylene, Ethylene (4);
6120-47 Electronegativity (1 + 7);
HM 2-10527 Formation of the Hydrogen Molecule;
2-10528 The Fluorine Molecule;
2-10529 Hybrid Orbitals;
2-10530 The Hydrogen Peroxide Molecule;
2-10531 The F_2O Molecule;
2-10532 The Water Molecule—Simple;
2-10533 The Water Molecule Hybridized;

2-10534 The Nitrogen Molecule;
2-10535 The Methane Molecule, CH_4;
2-10536 The Ammonia Molecule—Hybridized;
2-10537 Carbon Dioxide Structure;
2-10538 Structure of C_2H_4;
C 600-24 Oxidation States;
600-25 Oxidation States of Common Elements;
600-17 Ion Formation and Chemical Bonding;
600-45 Comparative Electronegativities of the Elements;
600-57 Resonance Bonding (Electron-dot Form);
600-58 Resonance Bonding (Conventional Dash Bond Form);
605-22 Electronic-Lewis Structures.

Programmed Instruction:

Barrow, G. M., *Understanding Chemistry II Chemical Bonding*, W. A. Benjamin, 1967

Cohen, M. H., *Beginning Chemistry*, Educators Publishing Service, 75 Moulton St., Cambridge, Mass. 02138

Dawson, C., *Chemistry, A Basic Systems Program—Atomic Structure and Bonding*, Appleton-Century-Crofts Lyons and Carnahan (1962)

Eichinger, J. W., *Chemicals Bonds/Introduction and Fundamentals*, Lyons and Carnahan (1968)

Kass, G. A., Individual Study Program in Chemistry, Vol. II, The Chemical Bond, W. A. Benjamin (1969).

Teaching Aids:

Charts

S-18806 Sargent-Welch, Table of Periodic Properties of the Elements

4844 Sargent-Welch, Relative Sizes of Atoms and Ions in the Periodic Table

Models and Model Construction: Refer to Chapter 5, p. 89

National Bureau of Standards Reference Data System, U.S. Dept. of Commerce Washington, D.C. 20234

Davent, B. deB., "Bond Dissociation Energies in Single Molecules," (1970).

TEACHER NOTES CHAPTER 7

Chapter Note: The activities for Sections 7.1–7.4 include the use of drill and exercise sheets and quizzes which are in the ANSWER SECTION for Chapter 7. Also note the teaching suggestions in PART II which are useful for the general course as well as a modified program.

Sections 7.1–7.2. Activities: Have students redraw a series of model diagrams as in Chapter 6 for NaCl, $CaCl_2$, $AlCl_3$, or $AlBr_3$. They should derive the representative formulas from the models of the component atoms. By writing the ionic formulas, i.e., $Ca^{++}Cl_2^-$ students readily recognize that the memorization of ionic charges makes it possible to write formulas without going through the drawing of models. Assign memorization of Table 7.1, schedule a quiz (the first of a series of quizzes needed for students to attain mastery through drill). Have students write the formulas in Sec. 7.2 with ions and charges given. Have them explain how they used the charges to obtain the formula. This leads to student statement of the principle(s) of the ion-charge method.

Section 7.3. Activities: Have students write a series of oxide formulas: Na_2O, CaO, Al_2O_3. From CaO they get the idea that where ionic charges are equal and opposite, no matter what their magnitude, no subscripts are needed. Develop the need for and placement of parentheses by using formulas in Sec. 7.3 and others such as $Al_2(SO_4)_3$, $(NH_4)_2SO_4$, $Ca(OH)_2$. Use **Laboratory Experiment 6:** "Combining Number," *Scientific Experiments in Chemistry*.

Table 7.2. Activities: Use Table 7.2 with students memorizing and using names from both systems. The drill sheet and quizzes include some examples. Define binary compound. Illustrate the two rules for naming binary compounds.

Section 7.5. Exhibits: Define and illustrate molecular formulas (from Chap. 6: H_2, O_2, N_2, F_2, Cl_2, HCl, H_2O, H_2O_2, NH_3, CH_4, C_2H_6, SO_2). Display their representative models.

Display crystal model of NaCl and models of other ionic crystals. Define empirical formula and cite other examples.

Define formula weight and illustrate calculation for H_2O and NaCl. Problems 1–7 deal with formula weight calculations.

Section 7.6. Activity: Define and illustrate molecular weight. Have students explain why formula weight is preferred by chemists. **Film:** CORF, *Molecular Weight of Oxygen.*

Section 7.6. Activity: Illustrate the calculation of the formula weight for compounds in Sec. 7-7. Have students try a more complicated example, $Cr_2(SO_4)_3$.

Sections 7.7–7.8. Laboratory Experiments: Experiment 9, "Determination of Oxygen in Potassium Chlorate"; Experiment 11, "The Formula of a Compound"; Scientific Experiments in Chemistry. Develop the steps used in finding the percentage composition of a compound and solve the test example (Fe_2O_3). Illustrate cancellation of Fe_2O_3 in the problem solution.

Section 7.8. Laboratory Experiments: Experiment 20, "Percentage of Water in a Hydrate"; Experiment 10, "Percentage of Water in Various Materials," *Scientific Experiments in Chemistry*. **Problems:** Solve the text problem on $Na_2CO_3 \cdot 10H_2O$, p. 128. Use problems 8–13. **Demonstrations:** Heat $CuSO_4 \cdot 5H_2O$, $KClO_3$ HgO or use films: YF-215, *Catalytic Decomposition of Potassium Chlorate*; YF-216, *Decomposition of Mercuric Oxide*; YF-219, *Electrolysis of Water.* You may

wish to use the $AgNO_3$ test for Cl^- ion in KCl and $KClO_3$ to complement the $KClO_3$ demonstration.

Section 7.9. **Activities:** Have students state the law of definite composition (Sec. 2.10). Have the law stated in terms of the percentage composition of the compound. Discuss the constancy of atomic weights as being basic to the atomic theory (Sec. 3.2, Sec. 3.13). Redefine atomic weight and illustrate if needed, the calculation of atomic weight from percentage of isotopes in naturally-occurring atoms (Cl, H, O). Use ball-and-stick models or chalk board diagrams of HCl (H_2O, CH_4, etc.). Relate the diagrams to the number of electrons shared, or transferred. Perform calculation for HCl as in text.

Section 7.10. **Experiment and Demonstration:** Wilhelm, D. L., "The Law of Definite Proportions," *Journal of Chemical Education*, 50, 436 (1973). Develop the law of multiple proportions as in Sec. 3.10 but also use parts by weight (atomic weights) for H_2O and H_2O_2 2:16 and 2:32 parts by weight respectively. **Film:** PSSC 0110, *Definite and Multiple Proportions.*

Section 7.11. **Review and illustrate:** the four quantitative definitions from Sec. 3.12 and Sec. 3.13: Avogadro number, mole, gram-atomic weight, atomic weight. **Activity:** See ANSWER Section Chapter 7 for a sample exercise sheet. **Exhibit:** In separate beakers have one mole of the following displayed: Hg(201 g), Zn(65 g clean, mossy Zn), S(256 g broken roll S), H_2O(18 g), $C_{12}H_{22}O_{11}$(342 g), C_2H_5OH(46 g), NaCl (58 g). **Define:** Have gram-molecular weight defined and illustrate it for diatomic gases and covalent compounds. Define and illustrate formula weight as representative for nonmolecular (ionic) compounds (as well as molecular compounds). Indicate that, for ionic compounds, the formula weight represents the mass of one mole and the Avogadro number of one mole of *formula units* of the ionic compound and the relevant number of moles of ions.

Section 7.12. **Demonstration:** You may wish to demonstrate the reduction of CuO by H_2 to provide student experience with procedures which may be used to provide data for the mass ratios of elements in a compound. See Experiment 9, "Hydrogen," *Scientific Experiments in Chemistry.* Use 0.5 to 1.0 g CuO in the reaction test tube. Unused hydrogen is burned at the jet exit. Before heating the CuO, the hydrogen issuing from the jet is ignited by taking collected test-tube samples from the jet, carrying the inverted test tube to the burner flame and igniting. When the hydrogen samples burn quietly at the mouth of the collecting test tube, the jet may be ignited by the flame of the burning sample.

Use the data from this demonstration or those for H_2O in the text to calculate the percentage composition of CuO or of H_2O. Then illustrate the calculation of the simplest formula by both methods, i.e., from percentage composition data and relative mass data to obtain the simplest mole ratio of atoms. Also use the Sample Problems. At the end of the experiment, blow out the jet and loosen the stopper in the generator tube to prevent possible flashback. Fuel gas may be used instead of H_2. **Problems:** Solve the text problems: (1) From percentage composition data, (2) from relative mass data. Solve the Sample Problems. Use problems 21, 22, 24, 25, 26–31.

Section 7.13. **Film:** MH 612008, *Determining a Molecular Formula.* **Develop the concept:** that for molecular substances: $(\text{empirical formula})_x = (\text{molecular formula})$, and $(\text{empirical formula weight})_x = (\text{molecular weight})$. **Transparencies:** As enrichment, you may wish to use one or more of the following: K & E 6120-28, "Molecular Weight Based on Depression of Freezing Point"; K & E 6120-29, "Molecular Weight Based on Boiling Point Rise"; K & E 6120-30, "Molecular Weight Based on Density." **Problems:** Solve the Sample Problem. Use problems 32, 33. Use Exercise 7, *Exercises and Experiments in Chemistry.*

Films:

PSSC 0110 *Definite and Multiple Proportions* (30 min.s);
MH 612008 *Determining a Molecular Formula* (16 mm, 12+ min, b/w,c,s);
CORF *Synthesis of a Compound* (16 mm, 13+ min, b/w,c,s). (Note: This is not too accurate.)
YF-253 *Model for Weight Relations in Chemical Reactions* (16 mm, 4 min,s);
YF-215 *Catalytic Decomposition of Potassium Chlorate* (16 mm, 4 min,s);
YS-216 *Decomposition of Mercuric Oxide* (16 mm, 4 min,s);
YF-219 *Electrolysis of Water* (16 mm, 4 min,s);
YF-214 *Catalytic Decomposition of Hydrogen Peroxide* (16 mm, 4 min,s);
YF-213 *Sodium Peroxide and Water* (16 mm, 4 min,s);
CORF *Molecular Weight of Oxygen* (16 mm, 11 min, b/w,c,s);
NF Grams-Atoms (8,c,s,2–5 min).

Filmstrips:

DG 643 The Mole Concept (fs, c, 40 frames)

Transparencies:

K & E 6120-35 The Mole of Nitrogen (6);
6120-48 The Law of Multiple Proportions (1+4);
6120-28 Molecular Weight Based on Depression of Freezing Point (1+5);
6120-29 Molecular Weight Based on Boiling Point Rise (1+5);
6120-30 Molecular Weight Based on Density (1+9);
C T 605-1 Formula Drill;
T 605-2 Valence Drill;
T 605-21 Chemical Formulas;
T 605-8 Units, Unit Cancellation;
T 605-10 Chemical Formulas, The Mole and Avogadro's Number.

TEACHER NOTES CHAPTER 8

Chapter Note: Activities and Demonstrations: Since this chapter deals with many chemical reactions it is necessary to provide a maximum of demonstrations, student experiments, and/or teaching aids which provide vicarious experience. Assigned homework equations and problems should be complemented by classroom reviews. The realism of demonstrations and experiments helps maintain student interest but ample provision for gradual development, drill, and written quiz must be made if students are to develop mastery of the necessary skills. Gradual development of the skills and understandings requires a carefully programmed use of the questions and problems in the chapter and complementary use of Exercise 8. Continuous reinforcement of formula writing skills is advisable.

Most students have had some previous experience with word and formula equations in elementary science. In Chapters 5 and 6 students have used equations related to ionization energy, electron affinity, oxidation-reduction, and bond energy. Practically all of these have had mathematical interpretation in terms of mole quantities. These experiences and their mathematical background may be appropriately integrated with the concept of an equality and the principle of conservation of mass (conservation of atoms).

Sections 8.1-8.2. Demonstration and Activities: Refer to Experiment 61, "The Preparation and Properties of Chlorine." Note Procedure 4 for caution in lighting the hydrogen jet. Form H_2O by inserting a burning H_2 jet into an inverted bell jar or large beaker. Have a student describe the demonstration. Ask for an expression of the chemical change in the fewest possible words. Write these on the chalkboard and develop the word equation explaining the meaning (word equivalents) of + and $\rightarrow$. Have students state what the word equation does and does not reveal about the chemical change. Have a student state the law of conservation of matter and energy. Develop the related law of conservation of atoms. Demonstration: Electrolysis of water (Film YF-219, *Electrolysis of Water*); Western slides: Decomposition and Synthesis of Water. Cite an example of a false equation (text) which is false because it is contrary to known facts.

Section 8.2. Review: It may be necessary to review Sections 7.5-7.8 on chemical formulas during some parts of Section 8.2. Laboratory Experiment 10: "Balancing Chemical Equations." Summarize: The three factors which must be considered in writing a balanced chemical equation.

Section 8.3. Activities: Apply the three factors to the text equation for the electrolysis of water. Use the abbreviations for the physical phases of reactants and products in the electrolysis of water equation and utilize where significant (volatility and precipitation) in subsequent equations. Define oxide. Review bonding and electronegativity related to oxides. Demonstrations: Burn sulfur in air (or oxygen), decomposition of mercury (II) oxide, zinc and hydrochloric acid, and aluminum sulfate and calcium hydroxide solutions. Demonstrate: Decomposition of HgO or use film YF-216 *Decomposition of Mercuric Oxide;* Zn + $HCl(H_2$ test); $Al_2(SO_4)_3$ + $Ca(OH)_2$ (observe precipitate). Refer students to Table 12 (Appendix: Solubility Chart). Develop the equations for the demonstrations as in the text to train students in the use of the steps in equation writing. Also write the equations for any of the word equations for Exp. 4, "Physical and Chemical Changes," Exp. 5, "Oxygen: Catalysis" that were not included in Section 8.3. Have students summarize the five statements at the end of Section 8.3. Laboratory Experiment 11: "Mass Relations in a Chemical Change." Laboratory Experiment 12: "Types of Chemical Reactions." Use results as the basis for the recognition and writing of the four general types of equations and derive the representative *general form* equation for each type. Demonstrations: from Experiment 12 and the others in Section 8.4 not in Experiment 12, i.e., Fe + $CuSO_4$, or Cl_2 + KI.

Sections 8.4-8.5. Demonstration: Have students identify and label gaseous product (g) by using Appendix Table 11, *Solubilities of Gases in Water.* They should be able to treat precipitates in a similar manner by using Table 12. Use these as an introduction to the idea of reactions going almost to completion. As a possible reason that an ionic reaction occurs, indicate the formation of a molecular species, i.e., H_2O, by demonstrating the solution of precipitated $Al(OH)_3$ in HCl followed by writing the equation. Demonstrate as many of the reactions as possible before developing the balanced equations. Identify as many of the products as possible of the reactions by physical or chemical properties, i.e., litmus for CO_2, NH_3, SO_2; limewater for CO_2; odor (waft) test NH_3, SO_2; $KMnO_4$ solution for SO_2, glowing splint test for O_2. Laboratory Experiment 9: "Determination of Oxygen in Potassium Chlorate."

Section 8.6. Demonstration: Demonstrate as many of the text examples as possible. In the replacement of a metal in a compound (aqueous salt solutions) use a variety of metals with various salt solutions as in Experiment 44, "Oxidation-Reduction Reactions," to establish relative positions of representative metals in activity series of metals (Table 8-1). Establish controls, i.e., Cu + $ZnSO_4$. Also use a variety of metals with water and acids. Use some that do not "work." To demonstrate that magnesium does react with hot water, heat Mg turnings in water or refer to Alyea, p. 142, "Magnesium burning in steam." To show that iron reacts with steam, place a moistened piece of asbestos sheet in the bottom of a Pyrex test tube supported horizontally. Iron filings are placed in the middle part of the tube and kept in place by either glass or steel wool. Place in the generating tube a one-hole stopper that has a delivery tube leading to a test

tube for gas collection by water displacement. The filings are heated by one burner and then a second burner heats the asbestos sheet converting the water to steam. Hydrogen is collected and identified by the "bark" test. Control experiments and "shakeout" solvents (CCl_4, CS_2, perchloroethylene which is replacing CCl_4) are used with the halogen replacement series. See Exp. 66, "Halide Ions" for details. **Transparency:** C T 600-1, "Activity Series." **Chart:** Welch-Sargent 4851, Electromotive Series. **Supplementary or complementary reference:** Fitzgerel, R. J., Kieffer, W. F. *Supplementary Readings for Chemical Bond Approach*, Chemical Education Publishing Co., 1960, "The Activity Series of the Metals," pp. 119-124.

Section 8.7. **Demonstration:** Reduction of Fe_2O_3 (rather than Fe_3O_4) may be demonstrated to complement the replacement of hydrogen in water (steam) by hot iron. The Fe product may be identified by its magnetic property and different color. Refer to burning H_2 and electrolysis of water as another example of reversibility. **Transparencies:** C T 600-55, "Equilibrium in Reversible Chemical Reactions." Western slides, "Decomposition and Synthesis of Water."

Section 8.10. **Review:** The quantitative interpretations of a chemical equation. Use with CO_2 equation. This is extended to the equation used in the text problem in Sec. 8.11.

Section 8.11. **Demonstration:** After establishing and illustrating the four steps in setting up the text mass-mass problem, illustrate the four operations leading to the final solution. Emphasize the cancellation of units and formulas in the solution operations. **Problems:** Solve the Sample Problems, noting that in Sample Problem 2, Part b does require the number of moles rather than the number of grams. Problems 1-5 deal with moles. Problems 6-10 are mass-mass problems.

Films:

MH EB *Chemistry: An Introductory Course* (on film), Encyclopaedia Britannica, (16 mm, 35+ min, b/w, c,s)

13 *Using the Mole Concept*

16 *Mole and Weight Relationships in Chemical Reactions*

For teachers who wish to review basics of chemistry of oxygen and related equations, the following series of short films is recommended.

YF-213 *Sodium Peroxide and Water* (2 min,s,c)

YF-214 *Catalytic Decomposition of Hydrogen Peroxide* (2 min,s,c)

YF-215 *Catalytic Decomposition of Potassium Chlorate* (3 min,s,c)

YF-216 *Decomposition of Mercuric Oxide* (3 min, s,c)

YF-217 *Distillation from Liquid Air* (2 min,s,c)

YF-218 *Condensation from Air* (1 min, s,c)

YF-219 *Electrolysis of Water* (3 min,s,c)

Transparencies:

K & E 6120-32 "Formula Weight" (1+5)

6120-58 "A Stoichiometric Problem" (1 + 7)

6120-59 "Balancing Equations by Inspection" (1 + 4)

C T 600-1 "Activity Series"

T 600-55 "Equilibrium in Reversible Chemical Reactions"

T 605-20 "Stoichiometry"

Color Slides (35 mm): Western Publishing Co.

W "Decomposition and Synthesis of Water"

W "Preparation and Properties of Hydrogen"

W "Preparation and Properties of Oxygen"

Filmstrips:

DG 643 The Mole Concept (40 f,c)

Loops:

WIL *On Writing Equations (I)* and *On Writing Equations (II)* (8 mm,s-8,4 min each, s)

Charts: Welch 4851 Electromotive Series Table 4841 Names, Formulas, and Charges of Some Common Ions. (This may be used for ready reference for formula writing.)

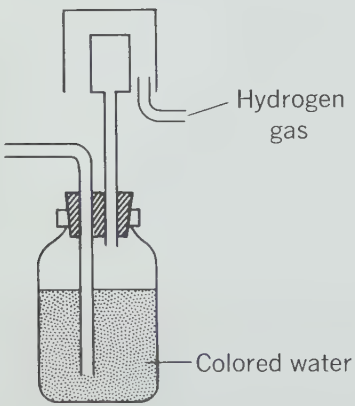
TEACHER NOTES CHAPTER 9

Chapter Note: **Activities and Demonstrations:** One or more of these devices may be used to introduce the kinetic theory since they provide a challenge calling for student observations and explanations. They may also be used at more specific sections in the text. (a) Drinking Duck (60264, Edmund Scientific Co.), (b) Cartesian Diver (1042, Welch Scientific Co.), (c) Simple Air Thermometer (77300, glass air-thermometer tube, Cenco). Construct by heating the bulb slightly and inserting the open end of the tube into a beaker containing colored liquid. Squeeze the bulb, by evaporation of water or alcohol and/or by placing a piece of ice on the bulb. (d) Molecular Vibration Tube (1724, Welch Scientific Co.); Molecular Demonstration Tube (77725, Cenco). Heat the mounted tube which contains mercury under reduced pressure and colored glass chips floating on it. Students may construct their own air thermometers or divers. A small pill bottle or vial for litmus paper may be made into a diver. Put it into a graduated cylinder and use either rubber sheeting or a stopper to close the cylinder.

Section 9.1. **Introductory (motivational) demonstrations:** "Drinking" duck, Cartesian diver, simple air thermometer, molecular vibration tube. **Review:** previous experiences indicating the particulate nature of substances; let two balls collide. Develop the three basic assumptions of the kinetic theory (Sec. 9.1). **Demonstration apparatus:** Molecular Motion Demonstrator, Educational Materials and Equipment Co.; Molecular Dynamics Simulator, Holt, Rinehart and Winston; Cenco 777-22, Kinetic Theory Apparatus.

Section 9.2. **Demonstrations:** Fuel gas collected by water displacement is opened. Odor is detected. Do waft test for opened bottle of household ammonia. Bring together moistened stoppers from HCl and $NH_3(aq)$. Note colored gases (NO_2, Br_2) escaping from opened container. Heat flask of air with attached balloon or manometer. Hold partially inflated balloon over hot steam radiator or hot plate (refer to HRW transparency CH-7126, Fig. 9-4). Heat Hg in Molecular Vibration tube (Welch 1724). Heat Iodine diffusion tubes (Welch 4425). Fill Brownian movement apparatus (Welch 4860) with smoke and view through microscope. Place a beaker containing H_2O + phenolphthalein and another containing concentrated $NH_3(aq)$ inside a large bell jar and observe the indicator change.

Section 9.3. **Demonstrations:** Drop a very small piece of Dry Ice into a flask with a balloon attached (molecular movement, volume and density of gas). Into one end of a wide glass tube in a horizontal position, insert cotton moistened HCl. Into the other end, simultaneously insert a cotton moistened with $NH_3(aq)$. The location of NH_4Cl indicates different rates of gas diffusion. Use the arrangement below with hydrogen or table fuel gas in the beaker above the porous cup to demonstrate rate of diffusion of low-density gas. Do not use bottled "country gas." Use chiefly methane in natural gas.



A 500- or 1000-ml flask, to which a balloon is attached, is heated. Before the heating, a student is asked to read the volume etched on the flask and is then asked, "What is the volume of the air in the flask?" The same question follows the heating. If the teacher desires, the flask may be heated by hot water enclosed in a large beaker. Several stoppered test tubes, one-hole stopper with long glass tube containing such liquids as water, mercury, and alcohol may also be heated by hot water so that the expansion of liquids and gases can be compared. The air in a stoppered can is heated so that the stopper is ejected, to show the effect of temperature on pressure. A compressed air popgun may be made from a foot-long section of a glass or plastic tube. One end is stoppered and a piston that fits the tube bore is inserted in the other end. Pushing down the piston causes the ejection of the cork. This shows the effect of decreased volume on pressure. A large bell jar, with a suitable base through which air may be withdrawn by means of a vacuum pump, may be used to show two effects of decreased pressure on volume. First a rubber diaphragm attached to the open end of the bell jar is pushed down into the jar as the pressure is decreased. (Call attention to the decreased air content inside the bell jar.) Second, when the pressure is decreased by the pump, the stopper on a bottle inside the bell jar is ejected because the air within the bottle is pushing against less pressure.

These demonstrations are on the general-science level, but they do help the student to resolve the problem. Why does the expression, "a cubic foot of air," mean little? The student realizes the need for selecting a standard temperature and pressure for gas-volume measurements.

Section 9.4. **Demonstration:** Have students explain the condensation of water vapor on the outside of a vessel containing ice. Define condensation temperature. Have students calculate the molecular weights of some of the gases and liquids in Table 9.1 and Sec. 9.4.

Define van der Waals forces. Have students calculate molecular weight of Ne (condensation temperature: -246°C). Compare with the molecular weights and condensation temperatures of H_2O and NH_3 to establish effect of molecular polarity (dipole-dipole attraction). Repeat for CH_4. Refer to the demonstration (Chap. 6) in which a thin stream of water is deflected by a charged rod. Summarize concepts in Sec. 9.4.

Section 9.5. **Exhibit and demonstrate:** The operation of barometer, manometer (J and U tube types), and pressure gauges. Connect a U type manometer to a thistle tube with a rubber diaphragm to show pressure effects. A mercury barometer inside the tall form bell jar (Welch 1483) on a pump plate is used to show the effect of reducing pressure. Heat a flask containing air and connected to a manometer (effect of temperature on volume and pressure of a gas). Exhibit a laboratory barometer. Refer to Fig. 9-8 and Fig. 9-7. Define standard temperature and standard pressure. Note: The pressure units, millimeters of mercury, torr, and atmosphere are to be abandoned eventually; and the SI pressure unit, pascal (newton per square centimeter), used exclusively, according to the recommendations of the General Conference on Weights and Measures and the National Bureau of Standards. Because the authors do not wish to use the unit, pascal, until it is more generally accepted by chemists, we are continuing to use millimeter of mercury because of its convenient association with barometric readings.

See *Journal of Chemical Education*, 48: 579(Sept. 1971).

Section 9.7. Laboratory Experiment 13: Gas Laws, precedes class work. Data for Procedure 1 is used for derivation of Boyle's Law. A form of Boyle's Law apparatus (Welch 1078, J-form or 1080 adjustable form) may be used in the classroom. *Project Physics Handbook*, Experiment 29 Behavior of Gases, describes the use of Elasticity of Gas apparatus (Cambosco 64-700) which may be used in the demonstration. Have students graph representative data (Exp. 13) or use the table of data below.

Pressure	Volume
200 mm	2000 ml
250	1600
333	1200
400	1000
500	800
800	500
1000	400
1200	333
1600	250
2000	200

Have students identify the parabola as representative of an inverse proportionality. Problems: Solve the text and Sample Problem (Sec. 9.8). In solving problems, have students use the reasoning process in approximating answers and the factor-label-cancellation process.

A demonstration of Charles' law which you may wish to use is described in the D. T. Haworth article on Charles' law, *Journal of Chemical Education*, 44, 353 (1967). State Fact 1 in Sec. 9.9. Use the data table below to derive fact 2 in Sec. 9.9.

Volume	Celsius	Kelvin
373 ml	100°	373°
323 ml	50°	323°
273 ml	0°	273°
223 ml	-50°	223°
173 ml	-100°	173°

The Kelvin scale should be based on the results of Procedure 2, Experiment 13 or on the data table (above). Transparency: HRW CH-7128 (Fig. 9-12). The Absolute Zero Demonstrator (Welch 1602) may be used to complement a redemonstration of a simple air thermometer and the effects of cooling and heating an enclosed air volume. This apparatus also called the Charles' law apparatus (Cenco 76408) shows the effect of increased temperature on the pressure of a gas having constant volume and mass (a variation of Charles' law). It may be used according to directions for Experiment 5-2, "Determination of Absolute Zero," *Laboratory Experiments for Foundations of Chemistry*. Solve the Sample Problem Sec. 9.11, preceding this activity by using Kelvin temperature conversion examples (see Table 9.2). Problems: 5-9.

Section 9.12. Activities: Derive the relationship for the combined gas laws. Solve the Sample Problem and problems 10-14.

Section 9.13 and 9.14. Demonstrate: Collection of an insoluble or slightly soluble gas by displacement of water. Attempt to collect a soluble gas such as NH_3 by displacement of water. Then collect a sample by displacement of Hg (use a test tube). Exhibit an eudiometer. Simulate the three situations in Fig. 9-13 by collecting an insoluble gas by displacement of water. Stopper collectors (2) and (3). Unstopper each successively in a tall cylinder of water. Manipulate each till the water levels inside and outside are equal (atmospheric pressure) and note the effects on the enclosed gas volume. Discuss the procedures for correction of volume and use Sample Problems for Sec. 9.13.

Discuss: the conditions under which real gases conform closely to ideal gas behavior and those under which they do not.

Films:

EBF 779 *Gas Laws and Their Application* (14 min, b/w);
CHEM 4106 *Gas Pressure and Molecular Collisions* (21 min, b/w, s),
CHEM 4115 *Molecular Motions* (13 min, c, s);
CHEM 4118 *Vibration of Molecules* (12 min, c, s);
MLA 0115 PSSC *Behavior of Gases* (15 min, s.);
YF-206 *Handling Gases;*
YF-222 *The Hydrogen Fountain* (3 min, c);
YF-250 *Vapor Pressure;*
YF-233 *Triple Point Determination;*
MH 612006 *Kinetic-Molecular Theory* (16mm, 8+ min, b/w, c, s);
EBF 2228 *Molecular Theory of Matter* (16mm, 11 min, b/w, c, s);
CORF *Laws of Gases* (16mm, 11 min, b/w, s);
CORF *Demonstrating the Gas Laws* (16mm, 21 min, b/w, c, s).

Loops:

Ealing Cartridged Super-8:
E 80-2967/1 *Properties of Gases* (3 min, 20 sec. c);
E 80-2959/1 *Diffusion* (3 min, c);
E 80-3379/1 *Diffusion Rates* (4 min, c);
E 80-3361/1 *Motion of a Molecule* (4 min, c);
E 80-3387/1 *Boyle's Law;*
E 80-2926/1 *Random Walk and Brownian Movement* (3.50 min, c);
ICF 13080 *Heat Affects the Pressure of Gases;*
ICF 13080 *Heat Expands Gases;*

ICF *Gas Pressure and Temperature;*
ICF 13220 *Brownian Movement* (8, 5-8, 3 min, c, s.);
ICF 13210 *Diffusion Explained* (8, s-8, 3 min, c, s.);
ICF 13250 *Diffusion Velocities* (8, s-8, 2 min, c, s);
ICF 13060 *Gases Condense as They Cool;*
LBF 15 *A Model of the Kinetic-Molecular Theory* (8, 4 min, c, s);
HR *Molecular Motion* (s-8, 4 min, c, s);
89-2810 *Kinetic Theory* (set of 6);
MLA 4001 *Absolute Zero* (s-8, b/w, s., 4 min).

Transparencies:
K & E 6120-1 Kinetic Molecular
Theory-Boyle's Law (2 + 2);
6120-2 Kinetic Molecular Theory — Charles' Law (3 + 3);
HM 2-10518 Boyle's Law Ideal Gas;
2-10519 Charles' Law;
2-10542 Molecular Distribution of Kinetic Energy;
C T 600-6 Boyle's Law (a);
T 600-7 Boyle's Law (b);
T 600-8 Charles' Law (a);
T 600-9 Charles' Law (b);
T 600-19 Kinetic Molecular Theory;
T 600-37 Temperature Scales;
T 600-62 Visible Evaporation of Mercury;
T 605-12 Boyle-Charles Combined Gas Law Formula;
T 605-14 Graham's Law of Diffusion;
McW Unit II The Gas Laws (8 in the Unit);
Boyle's Law
Charles' Law
The Kinetic Molecular Theory
MLA (Modern Learning Aids)
0916 Kinetic Theory of Gases

TEACHER NOTES CHAPTER 10

Section 10.1. Review: the gas laws and the similarities in the physical behavior of gases. Question: "Do gases exhibit any similarities in chemical behavior?" Loop and demonstration: Show the loop: E 84 0173 1, "Combining Volumes: Synthesis of Water." Demonstrate all or parts of the loop. (*Use a safety shield.*) In the electrolysis demonstration the two electrodes are placed in a thick, wide-mouth bottle which is wrapped in a wire shield. Eudiometers used in the synthesis contain platinum electrodes sealed in the closed end (Welch 5036-10 or 5036-20). Slides and film: use the Western slides, "Decomposition and Synthesis of Water" and the CHEM film 4103, *Gases and How They Combine.* The film provides the data for Gay-Lussac's law which leads to Avogadro's explanation.

Section 10.2. Activities: Refer to the CHEM loop, "Gas Pressure" (Chap. 9) in which the principles of gas pressure are related to Avogadro's law. Write the word equations for the combining volumes and then the corresponding number of molecules underneath for each of the examples (Sec. 10.2). Have students attempt explanation (Avogadro's) of the simple molecules of hydrogen, oxygen, chlorine, hydrogen chloride, water, and ammonia. Use molecular models to simulate reactions producing HCl. Have students use Avogadro's explanation for resolution of the Dalton-Gay-Lussac inconsistency. Use the transparency K & E 6120 33 "Gay-Lussac's Law of Combining Volumes and Avogadro's Principle." Discuss why Avogadro's explanation is now known as Avogadro's principle.

Section 10.3. Activities: Develop the supporting evidence for the diatomic composition of molecules of active gaseous elements showing that the simplest formulas must be checked by calculation of the molecular weights of gases from experimentally determined gas densities. Now write the formula equations underneath the sets of word equations for Sec. 10.2. Establish the significance of the coefficients and the relative volumes. Refer to student experiences in Experiment 10, "Balancing Chemical Equations," the answers to one of the related questions, deals with Avogadro's principle. Related laboratory experiments from *Scientific Experiments in Chemistry*, Experiment 12, "The Number of Atoms in a Molecule of Hydrogen," and Experiment 15, "Gay-Lussac's Law."

Section 10.4. Activity: Establish the monatomic nature of the noble gases. Use electron-dot symbols and the stability of the electron-pair (He) and octet (Ne, Ar) and the low condensations temperatures (Appendix, Table 16 Properties of Common Elements).

Section 10.5. Laboratory Experiment 16: "Molar Volume of a Gas;" review student results. If you do not let your students use $KClO_3$, perform as a demonstration. Students record data and make the calculations. An alternate experiment is described in *Chemistry*, 43; (1970) Jan., "Lab Bench-Molar Volume: A Safe Method." This experiment uses Dry Ice. Review the concepts of mole and Avogadro number and tabulate the related facts for molecular substances. Exhibit the 22.4-liter box (Welch 5262A Gram-Molecular Volume Cube). Some teachers have the formulas and gram-molecular weights of selected gases written on different sides of the box (Fig. 10-4) and state that the data has been experimentally determined. The equivalent HRW transparency is CH-7131. This leads to the definition of molar volume and the related generalization pertaining to the molar volume of all gases. Define gas density as the mass in grams per liter at STP and have students calculate the molar volumes of various gases (H_2, O_2, and others) from the gas densities. Exhibit a volumetric 1-liter flask. Focus attention on the functional relationships:

$$D(\text{of a gas}) = \frac{\text{g-mol wt}}{22.4 \text{ liters}} \quad \text{g-mol wt} = D \times 22.4 \text{ liters}.$$

Determine the density of SO_2 as in text Sec. 10.5. **Loop:** E 80-3288/1, "Identifying Gases by Density."

Section 10.6. Visuals: Use either the transparency set K & E 6120-30, "Molecular Weight Based on Density" or the Western slides, "Measurement of Molecular Weight." **Exhibit:** One of the variety of glass bulbs (Welch 4509A Balloon, Gas-Density-Dumas; Welch 4510 Glass Balloon with Stopcock) used in determining the density of gases. **Problems:** Solve the Sample Problem, reviewing concepts of partial pressure and correction of measured volume to STP. Students find the following to be a functional aid:

$$D(g) = \frac{g}{ml} \times \frac{1000\ ml}{liter}$$

which is followed by the use of g-mol wt = D × 22.4 liters. Solve the second Sample Problem dealing with conditions other than STP.

Section 10.7. and 10.8. Review: the use of the equation coefficient-mole relationships in gas reactions and its use in solving gas volume-gas volume problems. Apply the problem set-up of mass-mass problems for the text (Sec. 10.8) problem for H_2. **Problem:** Solve the Sample Problem including the procedure for calculating the volume of air.

Section 10.9. Problems: For the mass-gas volume problems (gas volumes given or required at STP), have the students suggest the two possibilities, i.e., volume given-mass required, mass given, volume required. Emphasize the mole relationships in the equation (Sec. 10.9) and the use of molar volume (22.4 liters) to determine the number of moles given in the typical problem before the use of operations 2-4 of mass-mass problems. A good technique is to use the results of the text problem, 17.9 g of $CaCO_3$ for having students devise the solution of a mass-gas problem with the mass given and the volume at STP (in this case, the volume of CO_2) required at STP. The reverse problem may be stated as: "How many liters of carbon dioxide at STP are produced by the decomposition of 17.9 grams of calcium carbonate?"

$$X = \frac{17.9\ g\ CaCO_3}{100\ g/mole} \times \frac{1\ mole\ CO_2}{1\ mole\ CaCO_3} \times 22.4\ liter/mole$$

$$= 4.00\ liters\ CO_2\ at\ STP.$$

In this type of problem, operations 1 and 2 of the mass-mass problem procedure are followed by multiplying the resultant expression (number of moles required) by 22.4 liters/mole. Follow up by assigning representative problems 18 and 19.

Section 10.10. Experiment: Have students perform Experiment 16, "Molecular Weight of a Gas" CO_2 or O_2, Experiment 5-5 "Molecular Mass of an Unknown Gas," *Laboratory Experiments for Foundations of Chemistry*. In these experiments the gas is collected at non-standard conditions and corrections to STP must be made. Develop the procedures for gases not measured at STP in Sec. 10.10 for (a) gas reactants and (b) gas products. Illustrate by using problems 30 and 31. In part (b) of problem 30, the gas volume calculated at STP from a given mass is converted to non-standard conditions. In problem 31, the mass is required and the gas volume is given at non-standard conditions.

Section 10.11. Problems: For mass-gas volume problems where the gases are collected by displacement of water, first review Sec. 9.14, Pressure of a gas collected by displacement of water, and apply to conversion of a gas volume at STP to nonstandard conditions involving collection by displacement of water. Solve the Sample Problem in Sec. 10.11.

Section 10.12. Review: The reasons for differences in the behavior of real gases versus ideal gas and the conditions for making measurements under which ideal gas behavior is approximated (Sec. 9.15). Derive the equation: $pV = nRT$ as in Sec. 10.12 by applying Avogadro's principle, Boyle's law, and Charles' law. Show how the units for the gas constant R are determined, i.e., dimensions of R being liter × atm per mole × °K. **Demonstrate:** Derive the numerical value of the gas constant, R = 0.082057 liter × atm/mole × °K for one mole of an ideal gas at exactly 1 atm and 273.15 °K. Apply the ideal gas equation to the equation to the text problem dealing with the determination of the molecular weight of a gas from gas density measurements, non-standard conditions. Use problems 41, 42, 43 for applications of the ideal gas equation. You may wish to illustrate how the ideal-gas equation simplifies the solution of (a) mass to gas volume problems and (b) a gas volume to mass problem. For (a), problem 30 is restated as: "What volume does the sulfur dioxide gas occupy at 25°C and 745 mm pressure when 50 grams of sulfur are burned? For (b) use problem 31. Compare the answers obtained previously. Transparencies which may be used include CT 605-26, "General Gas Law," K & E 6120-58, "A Stoichiometric Problem (1 + 7)."

Section 10.13. Discuss: the factors which contribute to the deviation of real gases from the perfect behavior of an ideal gas. List the molar volumes of ammonia, chlorine, oxygen, and nitrogen measured at normal conditions and their condensation temperatures: NH_3 (-33 °C), Cl_2 (-34.6 °C), O_2 (-183 °C), N_2 (-195.8 °C). Have students correlate molar volume deviations and condensation points. Review the related factors (See Sec. 9.4, Attractive forces between gas molecules) and show DG filmstrip 6509 van der Waals Forces.

Summary Review: You may also wish to use the frames on stoichiometry of the filmstrip DG 643, The Mole Concept, as a review of use of the mole concept in solving problems.

Films:
MH 612008, *Determining a Molecular Formula* (16 mm, 12 + min, b-w, c,s);
MLA 4103 (CHEM), *Gases and How They Combine* (16 mm, 8, 22 min, c, s);
CORF, *The Laws of Gases* (16 mm, 11 min, b-w, s);
EBF 779, *Gas Laws and Their Applications* (16 mm, 13 min, b-w, s);

Loops:
E 84-0173/1, "Combining Volumes-Synthesis of Water" (s-8, 3 min, c, silent);
E 80-3288/1, "Identifying Gases by Density" (s-8, 4 min, c, silent).

Filmstrips:
D G 643, "The Mole Concept" (40 frames, c);
D G 6509, "van der Waals Forces" (40 frames, c).

Transparencies:
HRW CH-7129, Molecules of hydrogen and chlorine are diatomic (Fig. 10-2);
HRW CH-7130, Molecules of elementary gas reactants are diatomic (Fig. 10-3);
HRW CH-7131, Molecular weight of gas (Fig. 10-4);
C T 605-26, General Gas Law;
K & E 6120-30, Molecular Weight Based on Density (1 + 9);
6120-33, Gay-Lussac's Law of Combining Volumes and Avogadro's Principle (6);
6120-58, A Stoichiometric Problem (1 + 7);

TEACHER NOTES CHAPTER 11

Section 11.1. **Demonstrations:** *Definite volume*—pour a measured volume of water (graduated cylinder) into a variety of containers, returning eventually to the graduated cylinder. This also shows fluidity, free surface, and variability of shape. *Diffusion*—layer alcohol over water or water over glycerine in a graduate by lowering a deflagration spoon just to the level of the denser liquid and pouring the second liquid into the bowl of the spoon through a thistle tube. Layers are visible because of differences in optical density. Liquids may be observed over several days. *Density*—have students lift (carefully) bottled samples of various liquids, preferably of the same volume; include mercury and iodoethane. *Attractive forces*—use the data in Table 11-1 and have students calculate the molecular weights of water, methane and benzene. Review the types of bonds, the role of dispersion interaction forces and that of dipole-dipole attraction forces. *Miscibility*—mix benzene and carbon tetrachloride, water and carbon tetrachloride. *Evaporation*—moisten chalk board with several liquids (water, ethanol, ether, perchlorethylene and a liquid Freon) simultaneously; note different rates of evaporation. *Odors*—waft test. **Laboratory Experiment 21:** Solution and Molecular Polarity; Brownian movement-carmine suspension viewed through microscope.

Section 11.3. **Demonstrations:** Dynamic equilibrium and vapor pressure. Insert several drops of liquids (water, ethyl alcohol, diethyl ether) into separate Torricelli barometers. Use a Florence flask connected to a manometer (colored liquid). Add several drops of ethanol or ether. Warm bottom of flask with hand or cautiously with a match, cool with an ice cube to demonstrate the factors involved in evaporation and condensation as well as the energy factors in displacement of equilibrium and establishment of a new equilibrium state.

Section 11.5. **Demonstration:** Develop statement of Le Chatelier's principle from vapor pressure demonstrations. Place a burning match next to the outside of the barometer. Alcohol (at the alcohol-mercury meniscus) shows increase in vapor pressure with temperature increase. To demonstrate the effect of increased volume on vapor pressure, use an evacuated flask with stopper attached to the Florence flask-manometer setup using a two-hole stopper. After equilibrium at constant temperature is reached, open the stopper in the evacuated flask and note the immediate and final effects on the manometer reading. Define equilibrium vapor pressure.

Section 11.6. **Activity:** Have water boiling in a beaker with thermometer inserted. Assign a student to record initial reading and readings at time intervals as boiling proceeds. Increase boiling rate. Keep temperature constant. To demonstrate water boiling at reduced pressure, stopper a flask of boiling water (a thermometer may be inserted into the liquid). Invert the flask, permit the water to cool. Produce boiling at a lower temperature by cooling the flask in running cold water. Air bubbles entering as stopper is loosened, shows reduced pressure. Use the HRW transparency CH-7133 (Fig. 11-6). Define boiling point. Use the HRW transparency CH-7134 (Fig. 11-7) as a source of data for student statements of the boiling points of the liquids, their boiling points at pressures below and above 760 mm pressure. **Reference and experiment:** Refer students to Williams, *Modern Physics*, pp. 200-201, for determination of heat of vaporization of water and to the related laboratory experiment for the determination of standard molar heat of vaporization.

Section 11.7. **Demonstration:** To demonstrate the liquefaction of gases such as ammonia or sulfur dioxide, cool the collecting tube in Dry Ice-acetone mixture. The heating effect of compression may be demonstrated by having a student use a hand pump to pump up a football or basketball. Relate liquefaction to kinetic theory attractive forces. Use Fig. 11-8. **Films:** YF-217, *Distillation from Liquid Air*; YF-218, *Condensation from Air*. **Define:** critical temperature, critical pressure. Use Table 11-2 to develop the two conditions needed to liquefy a gas and to explain the relationship of critical temperature to the nature of the attractive forces and molecular weights. Complement by exhibiting models and review molecular polarity.

Section 11.9. **Demonstration:** Formation of crystalline sulfur (rhombic by evaporation from solution of sulfur in CS_2, monoclinic by pouring molten sulfur into a folded filter paper, breaking top crust and pouring out the residual molten sulfur); amorphous sulfur by pouring boiling sulfur into cold water, removing the mass and showing its somewhat elastic property.

Sections 11.10 and 11.11. **Exhibit:** Display a beaker of melting ice cubes with inserted thermometer and take temperature at intervals. Heat a beaker of ice cubes (thermometer inserted) with the burner. Constant temperature. **Demonstrations:** Sublimation. Heat a few crystals of I_2 in an evaporating dish and collect recrystallized I_2 on the bottom of a water-cooled watch glass; expose a piece of Dry Ice(sublimation) or put a piece of it into a cylinder of water. Insert a small piece of naphthalene or paradichlorobenzene into the mercury column of a Torricelli barometer to show vapor pressure of a solid. To show softening property of amorphous solids: exhibit paraffin, heat it until melting, taking the temperature at intervals; heat a soft glass tube or rod until it softens and bends under its own weight (gradual softening instead of the sharply defined melting point of pure substances). **Demonstrate:** Iodine Diffusion Tubes (Welch 4425), Cryophorus (Freezing by Evaporation) Welch 1667; freeze the mercury in the Molecular Vibration Tube (Welch 1724). **Film:** YF-238, *Phase Demonstration* (Naphthalene). Refer to earlier Experiment 6, Energy and Entropy: Phase Change. **Reference, experiment, and demonstration:** Refer students to Williams, *Modern Physics*, pp. 197-199, for determination of heat of fusion of ice (molar heat of fusion) and the related laboratory experiment.

Sections 11.13 and 11.14. **Exhibit:** halite crystal and have students examine salt crystal with hand lens or microscope. **Exhibit:** other crystals, e.g., quartz, laboratory- or project-grown $CuSO_4 \quad 5H_2O$, alum, models of NaCl, Mg, Cu, diamond, graphite (commercial or constructed). **Demonstrate:** crystal formation by methods not demonstrated previously: cooling hot concentrated solution of NH_4Cl. Exhibit samples and models of the five types of substances in Table 11-3 and use or demonstrate properties and review binding forces (Sec. 11.14). **Laboratory Experiment 21** "The Solid State, Crystals, and Crystallization."

Sections 11.15 and 11.16: **Discussion:** Introduce the topic of water by referring to Fig. 11-15 and discuss related social and economic problems—pollution, water supply and population explosion, purification and recovery, desalination plants (including nuclear ones), drought, floods, etc. **Demonstration:** water purification by distillation or ion-exchange (Alyea: pp. 33, 87, 132, 219), (New York State Chemistry Handbook, pp. 123-129). **Demonstrate:** ice cubes floating in water (density). Repeat the demonstration showing water polarity —deflecting a stream from a buret with an electrostatically-charged rod. In conjunction with Table 11-4, use Fig. 8-8 and 8-9 from Williams, *Modern Physics* (density-temperature and volume-temperature changes for water). Explain the unusual properties of water in terms of hydrogen bonds and molecular structure. Use models and chalk board diagram.

Section 11.17. **Demonstrate:** the stability of the bonds in water by contrasting the difficulty of its thermal decomposition with the relative ease of decomposition of HgO. Use the student results of Experiment 18, "Chemical Properties of Water," Experiment 19, "Water of Hydration," Experiment 20, "Percentage of Water in a Hydrate," or use selected demonstrations from procedures in these experiments which are representative of the properties listed in Sec. 11.17. Write representative equations. Students may try Experiment 17, "Composition of a Mixture" (a mixture of two hydrates) from Weaver, *Scientific Experiments in Chemistry.*

Section 11.18. **Demonstration:** Early in a class period, place samples of sodium carbonate decahydrate, calcium chloride granules and pellets of sodium hydroxide into separate watch crystals so that the results are apparent when needed in considering efflorescence and deliquescence. Use the balance demonstrations in Secs. 11.18 and 11.19. Insert a small solid sample of the sodium carbonate decahydrate into the Hg column of a Torricelli or place in the equivalent vapor pressure set-up (flask and manometer). With the flask-manometer set-up, test effect on vapor pressure of the calcium chloride or sodium hydroxide. Another efflorescent substance, sodium sulfate decahydrate, is readily prepared from saturated solution by evaporation. Its long, large crystals can be stored until used.

Section 11.20. **Demonstration:** Use HRW transparency, CH-7107, "Hydrogen Isotopes." Draw atom-shell diagrams for the three isotopes of oxygen. Develop as in Sec. 11.20, 18 types of water molecules, production of D_2O by electrolysis, properties and uses of D_2O. **Student reports:** on Isotope Effects, pp. 332-333, Sienko and Plane, *Chemistry*, 4th ed.; Heavy Water, Quagliano and Vallarine, *Chemistry*, 3rd ed., pp. 204-205; list of *Scientific American* Offprints; list of articles from *Chemistry* and *Journal of Chemical Education.* **Film:** For summary on water, SUTH (MCA), *Chemistry of Water.*

Audio Visual Materials Films:

MH 611999, *The Structure of Water* (14 min,c,b/w, s).

YF-217 *Distillation from Liquid Air* (2 min,c,s);

YF-218 *Condensation from Air* (1 min,c,s);

YF-224 *Melting Points-Determination and Trends* (9 min,c,s);

YF-225 *Le Chatelier's Principle* (3 min,s,c);

YF-233 *Triple Point Determination* (15 min,s,c);

YF-234 *Movement in Crystal Growth* (5 min,s,c);

YF-250 *Vapor Pressure* (7 min,s,c);

YF-248 *Proton Motions in Ice* (5 min,s,c);

YF-238 *Phase Demonstration* (1 min,c,s);
YF-239 *Solution, Evaporation, and Crystallization* (3 min,c,s);
FAC (Collaborator: Lever Bros. Co.) *Physics and Chemistry of Water* (16 mm, 21 min,c,s);
EB 1675, *Explaining Matter, Molecules in Motion* (16 mm, 11 min,c,s);
4139 CHEM, *Crystals and Their Structures* (22 min, b/w,s);
4039 CHEM *Teacher Training Introduction to "Crystals and Their Structure"* (16 mm, 7 min, b/w,s), *Measurement of Vapor Pressure* (16 mm, 12 min,c,s);
0113 PSSC *Crystals* (16 mm, 25 min, b/w,s);
SUTH *Chemistry of Water* (16 mm, 14 min,c,s) MCA;
GE, *Our Air and Water* (16 mm, 15 min,c,s,) GE Educational Films, Building 705, Corporate Park, Scotia, N.Y. 12302;

Loops:
LBF LBFilms,
LBF 2 *Crystal Structures of Metals* (8 mm, 4 min, c);
LBF 16 *The Critical Point* (8 mm, 4 min,c);
EBF R80601, S80601 (8 mm, s-8, 4 min, c) Phase Change (Change of State);
HR 80-3403 *Boiling Point and Pressure* (s-8, 4 min, c),
80-2058 *Critical Temperature* (s-8, 4 min, c),
04-96331 *Close-Packing of Spheres* (s-8, 4 min, c),
04-96299 *Crystal Growth* (s-8, 4 min, c);
MLA 4006 *Crystals and X-ray Diffraction* (s-8, 4 min, b/w);
WD CHEM LOOPS 168 W 1005(T), 168 W 1025(K) *Crystals and X-ray Diffraction;*
WD 168 W 1009(T), 168 W 1029(K) *Molecule Motion in Condensed Phases;*
WD 168 W 1012(T), 168 W 1032(K) *A Problem in Diffraction;*
Bell System (local sources) *Crystals, An Introduction* (25 min,c,s);
STER (Free loan) *Krystallos* (Collaborator: Bell Telephone Labs.) (16 mm, 11 min,c,s),
Quartz Crystal Growing (Collaborator Western Electric (16 mm, 15+ min,c,s).

Filmstrips:
DG 629 *Crystals and Their Properties* (40 f,c),
DG 634 *Water, An Unusual Substance* (41 f,c),
DG 6502 *Le Chatelier's Principle* (40 f,c),
DG 6509 *van der Waals Forces* (40 f,c).

Loops:
EAL 82-0043/1 *Density of Liquids* (s-8, 2 min 30 sec, c),
82-0088/1 *Convection in Liquids* (s-8, 3 min, 15 sec),
80-3270/1 *Identifying Liquids by Density* (s-8, 4 min,c),
80-3262 *Identifying Solids by Density* (s-8, 4 min,c),
80-3304/1 *Thermal Expansion of Liquids* (s-8, 4 min, c),
80-3296/1 *Thermal Expansion of Solids* (s-8, 4 min,c),
80-0017/1 *Handling Solids and Liquids* (s-8, 4 min, c),
80-3403/1 *Boiling Point and Pressure* (s-8, 4 min,c),
80-3411/1 *Heat of Crystallization* (s-8, 4 min,c),
80-3429/1 *Heat of Fusion* (s-8, 4 min,c) Companion to 80-3411/1,
84-0116/1 *Bubble Model of a Crystal-Structure and Boundaries* (s-8, 2 min, b/w),
80-2058/1 *Critical Temperature* (s-8, 4+ min,c),
ICF 13220, 13225 *Brownian Movement* (8mm, s-8, 3 min,c),
13070, 13075 *Evaporation Causes Cooling* (8 mm, s-8, 3 min,c),
13090, 13095 *Boiling Points of Water* (8mm,s-8, 4 min,c),
13060, 13065 *Gases Condense As They Cool* (8mm, s-8, 4 min,c),
13030, 13025 *Liquids Evaporate* (8 mm,s-8,4 min,c),
13040, 13045 *Heat Expands Liquids* (8 mm,s-8,4 min,c),
13010, 13015 *Most Solids Melt* (8 mm,s-8,4 min,c);

Transparencies:
HRW CH-7132 Example of Physical Equilibrium and the Influence of Temperature (Fig. 11-4),
CH-7133 Standard Boiling Point of Water (Fig. 11-6),
CH-7134 Graph of Some Equilibrium Vapor Pressures (Fig. 11-7),
CH-7135 Unit Cells (Fig. 11-11),
CH-7136 Schematic Diagram of Six Basic Crystals (Fig. 11-12);
K & E 8120-4 Hydrogen Bonding in Water (1 + 4),
8120-5 Evidence for Hydrogen Bonding (2 + 4),
8120-30 Equilibrium Vapor Pressure (2 + 7),
8120-68 Change of State (1 + 5),
8120-69 Cooling Curve for Naphthalene (1 + 3),
8120-70 Eutectic (1 + 7);
HM 2-10533 The Water Molecule-Hybridized;
C T 600-20 Le Chatelier's Principle Applications,
T 600-49 Physical States of Matter;
McW Physical Properties Related to Bond Type,
The Effect of Hydrogen Bonding,
The Metallic Bond.

Slides:
W Measurement of Melting Point,
W Purification of Water.

TEACHER NOTES CHAPTER 12

Chapter Note: **Activities and demonstrations:** This chapter affords many opportunities for the application and extension of many of the concepts previously developed: bonding, phase changes, Le Chatelier's principle, Dalton's law of partial pressures, reaction

tendencies: energy and entropy. Effective use may therefore be made of models previously exhibited, as well as demonstrations, experiments, and other instructional aids that have been employed. The experiments which ideally are performed before classroom development are: Experiment 22, "Solubility, Rate of Solution, Heat of Solution"; Experiment 23, "A Solubility Curve"; earlier experiments—Experiment 20, "Solution and Molecular Polarity," Experiment 21, "The Solid State, Crystals and Crystallization"; and from *Scientific Experiments in Chemistry:* Experiment 20, "Making a Solubility Curve," Experiment 21, "Progressive Precipitation," and Experiment 22, "Finding Molecular Weight."

Section 12.1. Exhibit: variety of materials which emphasize the importance of solutions and suspensions: household ammonia, vinegar, and bleaches indicate percent of chemical; pancake syrup (sugar solution); tincture of iodine (solvent other than water). Samples of salad dressings and medicines (labeled "Shake well before using"), represent suspensions: Samples of water base paints and India ink are colloidal suspensions that may be diluted with water. **Demonstrations:** Make a sugar solution, dialysis of starch, glue, sugar, salt as in Sec. 12.1. Use demonstrations involving NaCl, $KMnO_4$ (visibility), MgO powder, clay, etc., so that students can recognize and state differences between solutions and suspensions, e.g., clarity, settling, filtration, homogeneity. Filter diluted India ink, then coagulate filtrate with acetic acid and refilter. Produce colloidal sulfur by adding HCl to a sodium thiosulfate solution in a tall cylinder. Use the beam from a flashlight (in a darkened room, if you wish), as the colloid forms from the thiosulfate solution, to illustrate the Tyndall effect (Fig. 12.1). Coagulate milk with acid. Demonstrations of dialysis (Fig. 12-2) and colloid phenomena: Alyea, pp. 49, 113, 146; Alyea, TOPS Demonstrations 259-267. Develop definitions in Sec. 12.1 **Demonstrations:** Demonstrate conductivity apparatus with water, salt and sugar solutions, glacial acetic acid, and diluted acetic acid. Exhibit selected examples (Table 12.1). Use Fig. 12-3. Use models to review bonding and polarity of H_2O, CCl_4, hydrocarbons. Review results of student Experiment 20, "Solution and Molecular Polarity." Demonstrate miscibility of ethanol and benzene (*use caution*), benzene + ethylene glycol, benzene + glycerol, mixtures of the alcohols + benzene with water. Other combinations that you may wish to try or assign to students for investigation include: solutes: iodine, acetic acid, paradichlorobenzene, sucrose, benzoic acid, potassium nitrate, sodium acetate; solvents: water, ethanol, benzene.

Section 12.1. **Demonstrations:** Miscibility: alcohol + water, glycerol + water, kerosene + benzene, kerosene and CCl_4 (Care). **Laboratory Experiment 20:** Solution and Molecular Polarity.

Section 12.4. Review: Hydrogen bonds (Sec. 11.-16). Relate to abnormally high boiling point and melting point (ice). See Fig. 12-5. Relate to solubility of ethanol in water. Use chalk board formulas for H_2O: long dashes for covalent bonds and a series of dots representing hydrogen bonds between molecules (See Fig. 9-20), Toon and Ellis, *Foundations of Chemistry.* **Demonstration:** "Hydrogen Bonding in Liquids," Alyea, p. 143.

Section 12.5. Activities: Start with a very small amount of a very soluble substance (NH_4Cl, KNO_3), gradually add more and more solute until a saturated solution results. Define saturated solution and solution equilibrium. Exhibit equivalent of Fig. 12-6. Although some of these terms do not appear until Sec. 12.11, students can identify the solutions produced as dilute, concentrated, and eventually saturated. Have them identify the first two terms as also being unsaturated and attempt pairings such as dilute unsaturated, concentrated unsaturated, concentrated saturated. Make up a saturated solution of boric acid in cold water as an example of a dilute, saturated solution. Shake up a few iodine crystals with water and heat to show increase in solubility upon heating and evidence of a *dilute* saturated solution. Save the saturated solution of NH_4-Cl for later use in demonstrating the effect of temperature changes on solubility. Add H_2O to a saturated solution of sugar, shake and note effect on amount of undissolved solid. Decant a portion of the liquid (saturated solution) and have students explain how they could experimentally distinguish between this solution and an unsaturated solution. Use Fig. 12-7 and weigh out the solutes as shown. **Film:** MH-612005, *Dynamics of Solution.* You may wish to set up a saturated solution of NaCl (small crystals) in a stoppered vessel and eventually large cubic crystals grow as surface area minimizes. The change is faster at elevated temperatures. This may be a simple worthwhile project for students, with quantitative aspects, i.e., find the weight of original undissolved solid, find weight of solid left undissolved (both at room temperature).

Section 12.6. **Demonstrations:** Pressure and Solubility: uncap a bottle of carbonated beverage. Pour it into a beaker. Students note and explain effervescence. Place a sample in an open container on the base of a vacuum table, cover with a bell jar, and reduce the pressure. A water aspirator may be used instead of the vacuum table and vacuum. Test (with limewater) a sample of soda water which has been standing exposed to air for some time and a sample which has stopped effervescing. Demonstrate an NH_3 or HCl fountain (Alyea, p. 13) as an example of compound formation. State Henry's law. **Discussion:** Pressure changes and solubility in terms of Le Chatelier's principle. Review Dalton's law of partial pressures (Sec. 9.14).

Section 12.7. **Demonstrations:** Let a tall cylinder of cold tap water stand from the beginning of the

period. Have students note gas bubbles as the solution warms. Hold a test tube containing tap water almost horizontally, heat slowly and gently. Have students note gas bubbles which appear well before boiling starts. Use the data in Table 10, Appendix, for the effect of temperature on the solubility of gases. Warm the saturated solution of NH_4Cl and discuss equilibrium disturbance. After solution of the undissolved solute is complete, cool the hot concentrated solution under the tap. Have students note crystallization. Use HRW transparencies: CH-7137 (Fig. 12-8), CH-7138 (Fig. 12-9), and data in Table 12.2 to have students arrive at generalization about effect of temperature on solubility of solids in liquids. Have them identify substances that do not conform. Discuss the generalization as it resembles phase changes and the energies involved. See Alyea, p. 189, "The Effect of Temperature on solubility," which involves the formation and solubilities of various hydrates of $MnSO_4$. See the *Journal of Chemical Education*, 47,275 (1970) for demonstration using calcium acetate. Its saturated solution yields a precipitate upon heating (8 g of solid in 25 ml H_2O). **Laboratory Experiments**: 22, "Solubility, Rate of Solution, Heat of Solution"; 23, "A Solubility Curve"; or Experiment 20 from *Scientific Experiments in Chemistry*.

Section 12.8. **Demonstrations**: Increasing the rate of solution: use student results of Experiment 22. Demonstrate effects of stirring, powdering, and heating for $CuSO_4 \cdot 5H_2O$; use equal masses of cube, granulated, and powdered sugar to study rate of solution and surface area.

Section 12.9. **Demonstrations**: Dissolve NH_4Cl, KNO_3, NaOH, and H_2SO_4 separately in water and have students note temperature changes by touch and thermometer readings. Dissolve a gas (air-cooled) such as NH_3 or HCl in water at room temperature. Take the temperature of resultant solution. Relate the findings of these demonstrations to Table 12.3. Use (a) a plastic-ball model (Fig. 12-10), and (b) a lattice or crystal model of NaCl and ball-stick models of H_2O to develop the three important actions that occur during the solution process (Sec. 12.9). For (b), simulate the solution of Na^+ by attachment of the negative ends of the polar H_2O molecules, of Cl^- by the positive ends of H_2O molecules. For (a), use student results of Experiment 20, "Solution and Molecular Polarity," and discuss types of forces of molecular attraction including hydrogen bonding. Discuss the nature of the energy changes for the three actions, the entropy changes, and the final demonstrated solubility. This discussion should be in terms of the net energy change versus entropy change for (a) solids, (b) gases. Use Question 20. Use the general and specific equations in Sec. 12.10. Discuss the meaning of the positive and negative values for heats of solutions (Table 12.3). Explain the effect of temperature changes on the solubility of solids, gases, and liquids. The explanations should be made as applications of Le Chatelier's principle. Relate data in Tables 12.2 and 12.3. **Optional demonstrations**: Dehydrate Epsom salts, Insert a thermometer into the dry $MgSO_4$ and add water, several drops at a time. Smith, W. L., *Chemistry*, 45, (Apr. 1972), "Thermochemistry and First Aid" utilizes instant hot packs and cold packs. The demonstration is relevant and utilizes instant hot packs and instant cold packs, has value in showing relationships of heat (enthalpy) of solution and heat (enthalpy) of hydration as well as applying Le Chatelier's principle to solubilities of salts during temperature changes. The cold pack contains the dry chemical NH_4NO_3 and a pouch of water. The hot pack contains dry calcium chloride and a pouch of water. Striking the packs causes the pouch to break. The pertinent data are NH_4NO_3: solubility at 20°C is 192 g per 100 ml of water,

$\Delta H_{soln} = +77$ cal per gram

$CaCl_2$: solubility at 20°C is 74.5 g per 100 ml H_2O,

$\Delta H_{soln} = -178$ cal per gram

Section 12.11. **Review or redemonstrate**: Qualitative concentration terms. Demonstrate dilution by addition of water, and concentration by evaporation or addition of solute. Exhibit or weigh out masses representing one mole of selected solutes. Prepare representative solutions of varying molalities. Label these with the specific *m* concentrations.

Section 12.12. **Demonstration and Activities**: Recall or redemonstrate vapor pressure of water (and effect of temperature change) using a Torricelli barometer or equivalent apparatus. Insert a drop of salt solution into another barometer tube. Use HRW transparency CH-7141 (Fig. 12-11) to explore details of effect of solute on vapor pressure and later on freezing and boiling points. Discuss the uses of antifreeze and coolants in automobiles. Have students inquire at local service stations and garages about the amounts of these substances recommended for different automobiles (varying radiator capacities). Exhibit and use the hydrometer used for testing concentration of antifreezes. Note effect of temperature on recommended concentration (amount of antifreeze). Compare rates of melting of ice cubes or crushed ice with and without addition of solid NaCl (rock salt). Base the development of K_f on Sec. 12.12, and the related data in Table 12.5. **Problems**: Solve the Sample Problems in Sec. 12.12. Use problems 10, 12, 13.

Section 12.13. **Transparency and Demonstration**: HRW CH-7141 for boiling point elevations for nonvolatile nonelectrolytes. To demonstrate elevation of boiling point, take the temperatures of boiling solutions of (a) sugar, (b) salt. Use Table 12.5 for K_b. **Problems**: 11, 14.

Section 12.14. **Discussion**: Develop the relation-

ships leading to the expression for solving for g-mol-wt. Solve the Sample Problem in Sec. 12.14. Experiment, activities, and problems: Interested students may perform Experiment 22, "Finding Molecular Weights (Masses)," *Scientific Experiments in Chemistry*. Others may wish to determine the molecular weight of urea, using 8 g of urea in 40 ml of water. Measure the freezing point depression by placing the test tube into a salt-ice mixture in a thermocup. See the *Journal of Chemical Education*, 47, 275 (1970). Solve problems 15-17.

Films and Programmed Instruction:

Films:

EBF 201 *Colloids* (16 mm, 11 min, b/w, s);
CORF *Principles of Ionization* (16 mm, 13+ min, b/w,c,s);
LGC E-66 & E 66A *The Properties of a Covalently Bonded Molecule; The Structure of a Covalent Molecule, CCl_4* (8 mm, 3 min,c,s);
SUTH *Solutions* (8 mm, 3 min, c,s);
UNESCO 2, 3 *Volume Changes on Mixing Liquids I & II* (8 mm, 4 min,c,s),
7 *Solid Solutions* (8 mm, 4 min,c,s);
AIM YF-239 *Solution, Evaporation, Crystallization* (16 mm, 3 min,c,s);
ICF 13220, 13225 *Brownian Motion Models* (8 mm, s-8, 3 min, c,s);
FAC *Physics and Chemistry of Water* (16 mm, 21 min, c,s);
MH 612005 or 612017 *Dynamics of Solution* (16 mm, 14 min,b/w,c,s);
612002 or 612014 *Liquids in Solution* (16 mm, 11 min, b/w,c,s);
611999 or 612024 *The Structure of Water* (16 mm, 14 min,b/w,c,s);
CORF *Solutions* (16 mm, 16 min,b/w,c,s);
Properties of Solutions (16 mm, 28 min, b/w,s);
Molecular Weight of Solutes (16 mm, 8 min,b/w, c,s);
The Colloidal State (16 mm, 16 min,b/w,c,s);
MLA CHEM 4154 *Shapes and Polarities of Molecules* (16 mm, 18 min,c,s);
E 80-3429/1 *Heat of Fusion* (s-8 cartridge, 4 min, c,s);
E 80-3411/1 *Heat of Crystallization* (s-8 cartridge, c,s);

Filmstrips:

DG 634 *Water—A Most Unusual Substance* (40 f,c),
DG 661 *The Solution Process* (40 f,c),
DG 665 *Colloids* (40 f,c).

Transparencies:

HRW CH-7137 Solubility curves (Fig. 12-8)
CH-7138 Solubility plotted as a function of temperature (Fig. 12-9)
CH-7140 Molecular diagrams of three common solvents (Fig. 12-4)
CH-7141 Vapor pressure plotted as a function of temperature for a pure and dilute solvent. (Fig. 12-11).
K & E 6120-4 Hydrogen Bonding in Water (1 + 4)
6120-5 Evidence for Hydrogen Bonding (2 + 4)
6120-6 Dissolving Sodium Chloride in Water (3 + 1)
6120-14 Ionization Reaction in Water-Liquid Solution (2 + 1)
6120-15 Ionization Reactions in Water-Gas (2 + 2)
6120-16 Hybridization (1 + 5)
6120-28 Molecular Weight Based on Depression of Freezing Point (1 + 5)
6120-29 Molecular Weight Based on Boiling Point Rise (1 + 9)
6120-40 Equilibrium in Saturated Solution (1 + 2)
6120-41 Concentrations of Solutions (1 + 4)
HM 2-10533 The Water Molecule-Hybridized
2-10533 The Water Molecule-Hybridized
2-10545 Aqueous Solubilities — I
2-10546 Aqueous Solubilities — II
C T600-10 Dissolving of Ionic Salts by Water
T 600-16 Hydration of Ions in Water Solution
T 600-20 Le Chatelier's Principle Applications
T 600-36 Solubilities of Various Substances
T 600-43 Solubility-Factors Affecting the Extent of Solubility of Substances in Liquids
T 600-44 Semi-quantitative Solubilities of Compounds
T 605-3 Solubility and Activity Drill
T 605-5,6 Electrolytes
T 605-11 Dalton's law of Partial Pressures
T 605-19 Molecular Weight Determination from Freezing Point Data
T 600-27,28 Solutions
McW Dalton's Law of Partial Pressures
Solution of an Ionic Salt
Solubility Rules
Solubility Curves
The Effect of Hydrogen Bonding

TEACHER NOTES CHAPTER 13

Chapter Notes: Activities and demonstrations: Many of the demonstrations in this chapter may be repeats of those used in previous chapters (Chapters 6 and 12). The emphasis is different. Where demonstrations of conductivity of molten electrovalent compounds and of solutions were previously used to establish some reality upon which bonding concepts were based, in this chapter, bonding concepts are used to explain the demonstrated reality. Extensive use is again made of models to complement the demonstrations and help develop equation writing skills. Although behavioral objectives are stated separately, some at least, such as those relating to electrical conductivity and colligative properties are dealt with almost simultaneously as the fact-gathering process leads to the inductive aspects of concept development.

The chapter furnishes an excellent opportunity for the development and understanding of scientific methods and scientific attitudes. One good approach is to start with Experiment 25, "Conductivity of Solutions," from *Laboratory Experiments in Chemistry*, as a teacher demonstration. The experiment as described lends itself to almost infinite variation. The experimental approach should allow the class to recognize phenomena which call for explanation. First, students should be challenged to offer their explanations of why electrolytes conduct and nonelectrolytes do not. Student theories should be explored. For example, students may offer the Faraday explanation. This may then be examined in the light of other evidence such as freezing point and boiling point phenomena, in which electrical potential (the student may say electricity or electric current or electric voltage) plays no part. Some students may offer explanations similar to the Arrhenius theory. These should be checked against previously developed concepts of bonding and structure. Some students will undoubtedly observe that in the demonstration some solutions caused a bright light while others caused a dim light. The story of the initial reception of the Arrhenius theory by his chemistry professor should be told at an appropriate point. An attempt to evaluate the professor's reaction should also be made. According to one version of the story, Arrhenius came to his conclusions during the night of May 17, 1883, after two years of ceaseless toil. The next morning he rushed to his chemistry professor and said, "I have a new theory of electrical conductivity." The professor looked at Arrhenius rather disdainfully and replied, "How interesting. Goodby." The class may then be asked to evaluate the professor's reaction as a scientific attitude. According to one version, the professor explained his reason for failing to listen. He had heard many theories of ionization, all of which were soon proved wrong or inadequate. On the basis of previous experience, he reasoned that this was merely another unacceptable student theory.

Sections 13.1 and 13.4. Demonstrations: Use Alyea page 15, demonstrations 6-1, 6-2, "Conductivity of Solutions." Also Alyea pp. 74–75, 6-1s to 6-7s. (Conductivity of Solutions): Start with observation by the class of a simplified, plug-in conductivity apparatus. Draw the diagram on the chalk board. After asking students to explain why the lamp does not light (use an inserted switch and close it), put first the wooden handle of a knife, then the blade across the gap between the electrodes. Students realize that some solids (metals) are conductors and others are not conductors. Generally they will explain metallic conductivity as being due to movement of electrons (emphasize that they are *charged particles*) through the solid conductor. Put a large halite crystal across the gap to show that a solid salt does not conduct. Exhibit a crystal model of NaCl, review electrovalent compounds, have students identify the ions and write the formula $Na^+Cl^-(s)$. Have students explain that conduction does not occur because there are no free electrons and that the ions are fixed in the NaCl lattice and cannot move. At this point you may wish to insert an ice cube (held in rubber gloves or rubber-tipped tongs) across the electrode gap to show that solid water is also a nonconductor. Granular NaCl in a beaker or crucible is similarly tested. Fuse the NaCl by preheating a crucible with a high-temperature burner, and dropping in half a teaspoon of granular NaCl. Have students recall the general high melting point of electrovalent compounds (salts). After fusion, drop in successive small portions of NaCl until the melted salt covers the bottom of the crucible. Insert the electrodes and have students explain conductivity in terms of *free-moving* ions. Distilled water is tests for conductivity. It is added to NaCl(s) in a beaker, stirred, and the solution is tested for conductivity. The question as to which substance, NaCl or H_2O, is conducting is raised and resolved. The conclusion reached is that the ions of NaCl become free-moving in water solution. Two representative equations for dissociation of electrovalent compounds are written:

$$\underset{\text{(fixed in lattice)}}{Na^+Cl^-(s)} \longrightarrow \underset{\text{(free moving)}}{Na^+ + Cl^-}$$

$$\underset{\text{(fixed in lattice)}}{Na^+Cl^-(s) + H_2O} \longrightarrow \underset{\text{(free moving)}}{Na^+(aq) + Cl^-(aq)}$$

Review freezing and boiling point variations for nonelectrolytes (Secs. 12.12, 12.13). Provide the data for electrolytes in Sec. 13.1. Demonstrations and experiments: Boiling point elevation for 1 *m* NaCl, $CaCl_2$, $C_{12}H_{22}O_{11}$, using a Beckman thermometer (Welch 5698); Experiment 22, "Finding Molecular Weights," *Scientific Experiments in Chemistry*; Experiment 31, "Molecular Weight Determinations," *Chemical Principles and Experiments*. Have students evaluate Faraday's explanation in the light of freezing and boiling point data. Use the caption for Fig. 13-2 or the Arrhenius anecdote in the introductory suggestions. Have student state Arrhenius' explanation and evaluate it in terms of present-day knowledge of crystalline electrovalent compounds and the polar nature of water.

Section 13.5. Activities: Refer to dissociation equations or, if not previously written, develop as in Sec. 13.4. Use Figs. 13-3, 13-4, 13-5, 13-6, the available transparencies for Fig. 13-3, and Fig. 13-6. Make a saturated solution of NaCl (40 g) in 100 ml of H_2O; write the equation using $\rightleftarrows$ for the ionic equilibrium. Weigh out 80 g of $NaNO_3$ and dissolve to emphasize the difference in solubility. Add more $NaNO_3$ to saturation and write the equilibrium equation. Add a very small amount of $BaSO_4$ to water and stir to emphasize practical insolubility. Precipitate AgCl ($KCl + AgNO_3$) and have students identify the precipitate by using Table 12 in the Appendix, Solubility of Compounds. Write the empirical, ionic, and net ionic equations.

Section 13.6. Exhibit: a bottle of glacial acetic acid and have students predict conductivity. Test for conductivity, dilute, and retest to show ionization. Repeat conductivity tests for benzene, HCl in benzene, and then add water to the latter, stir, and test again. Using models, reestablish the polar nature of HCl and H_2O. Have students explain how ions can be produced by ionization due to dipole-dipole interaction and hydrogen bonding: HRW transparency CH-7144 (Fig. 13-7). Demonstration: Add $AgNO_3$ solution to the water solution of HCl (benzene-HCl), thus identifying the chloride ion present in that solution. Add blue litmus and establish the existence and this property of the hydronium ion. Exhibit a spatial model of the simple hydronium ion. Also use the transparency HRW CH-7145 (Fig. 13-11) for the hydronium ion. Use the marginal note for Sec. 13.8 for $H_9O_4^+$. Demonstration: At this time you may also write the equation for ionization of NH_3. Demonstrate the nonconductivity of its benzene solution and, after adding water, the conductivity of its water solution. Use models, diagrams, and equations similar to that for Fig. 13-7. Identify the OH^- ion by litmus and contrast the dissociation of metallic hydroxides (soluble) with the ionization of NH_3.

Section 13.7. Demonstration: Redemonstrate the intensity of the light bulb for HCl versus acetic and boric acids; NaOH versus NH_3(aq). Try the series NaOH, $Ca(OH)_2$, $Ba(OH)_2$, and $Mg(OH)_2$, using filtered solutions of the last three. Lecture-table meters may be used to show that solutions which have little, if any, effect on the light bulb, do have perceptible conductivity. Salts which dissociate to only a slight extent by conductivity evidence are $HgCl_2$, $CdSO_4$.

Section 13.8. Demonstration: First test the conductivity of H_2O with a 25 watt bulb. Then use either a neon or an argon lamp. Have students suggest a mechanism for ionization. Use the HRW transparency CH-7145 (Fig. 13-11), complement with models, formulas, and equations. Discuss the role of the hydrogen bond. Refer students back to Fig. 6-11 and Section 6.17 (page 115) along with Sections 11.6 (page 205) and 12.4 (page 226) if necessary. Add red and blue litmus paper to water to establish the equivalency of the hydronium and hydroxide ion concentrations and the consequent acid-base neutrality of water.

Section 13.9. Demonstration: Demonstrate the nonpolar nature of CCl_4 (or substitute benzene) by the attempted deflection of a thin stream from a buret by a charged rod. Review nature of covalent bonds, and spatial arrangement of polar bonds (models and electronegativity difference). Demonstrate insolubility in water. Review the dissolving of sucrose in water, Fig. 12-10. Test for conductivity of its solution and discuss freezing and boiling point effects (See Fig. 13-12a).

Section 13.10. Discussion: refer to Sec. 12.4. Explain freezing point and boiling point effects, if not already considered, by the number of particles (ions) produced per mole in dilute solutions. As other than text examples, write equations for NaCl, Na_2SO_4, $Al_2(SO_4)_3$ and have students tell the total number of moles of ions produced by complete dissociation from one mole of solute.

Section 13.11. Discussion: degree of ionization–apparent degree of ionization: base the discussion on Table 13.1. Use the marginal note for Sec. 13.11 and the HRW transparency CH-7146 (Fig. 13-13). A useful table from *Handbook of Chemistry and Physics* is:

Approximate degree of ionization for active salts (0.1-N solution)

Type R^+ R^- (e.g. KCl)	0.86
Type R^{++} $(R^-)_2$ (e.g. $BaCl_2$)	0.72
Type $(R^+)_2$ R^{--} (e.g. K_2SO_4)	0.72
Type R^{++} R^{--} (e.g. $CuSO_4$)	0.45

Films

CORF *Ionization* (16 mm, 18+ min, b/w, c, s);
CORF *Principles of Ionization* (16 mm, 13+ min, b/w, c, s);
INDU *Ionization and Ionic Equilibrium* (16 mm, 15 min, b/w, c, s);
MLA CHEM 4109 *Electric Interactions in Chemistry* (16 mm, 21 min, c, s).

Loops

MLA CHEM LOOP or WARD 168 W 1008(T) or 168 W 1028(K) *Ion Exchange* (s-8, 4 min, c);
HRW *Ionic Reactions in Aqueous Solutions* (set of 10, s-8, c).

Transparencies

HRW CH-7142 Cubic structure of crystalline Sodium Chloride (Fig. 13-3);
HRW CH-7143 Hydration of ions (Fig. 13-6);
HRW CH-7144 Ionization of polar Hydrogen Chloride (Fig. 13-7);
HRW CH-7145 Formation of hydrogen bond between two water dipoles (Fig. 13-11);
HRW CH-7146 The apparent number of ions present may be less than the actual number (Fig. 13-13);
HRW CH-7141 Vapor pressure plotted as a function of temperature for a pure and dilute solvent;
K & E 6120-6 Dissolving Sodium Chloride in Water;
K & E 6120-14 Ionization Reactions in Water—Liquid Solution;
K & E 6120-15 Ionization Reactions in Water—Gas;
C T 605, 606 Electrolytes;
C T 600-10 Dissolving of Ionic Salts in Water;
C T 600-16 Hydration of Ions in Water Solution;
C T 600-17 Ion Formation and Chemical Bonding;
C T 605-18 Effect of Ionization on Colligative Properties;
McW Ion Formation.

Filmstrips

DG 620 Electrochemistry—Linking of Two Sciences
MH Ionization
MH Electrolysis

References

Alyea, *Tested Demonstration in Chemistry*.

Bell, *The Modern Theory of Electrolytes*, Murray, London.

Castka, "Modern Chemical Concepts and the High School Curriculum," *Journal of Chemical Education*, 17, 487 (1940).

Jaffee, B., *Crucibles*, Simon and Schuster, 1948.

Page, J. A., et. al., *Chemistry Principles and Experiments*, Holt, Rinehart and Winston, 1969.

TEACHER NOTES CHAPTER 14

Activities and demonstrations: The chapter treatment calls for the application of ideas of bonding and ionization previously developed in chapters 6 and 13. It provides opportunities for the review and strengthening of such ideas through the use of electron-dot formulas, models, oxidation states, and electronegativity. The experiments which students can perform before appropriate portions of the chapter are: Exp. 24A, "Preparation and Properties of Acids," and Exp. 24B, "Preparation and Properties of Hydroxides." Related experiments from *Scientific Experiments in Chemistry* are Exp. 24, "Hydrogen Chloride, Hydrochloric Acid;" Exp. 25, "Reactions between Oxides and Water;" Exp. 26, "Conductivity" (strength, concentration, and quantity of acids); Exp. 27, "Acid and Alkaline Solutions;" Exp. 28, "Rate of Reaction as Determined by Strong and Weak Acids;" Exp. 29, "Preparation of Soluble Salts." You may wish to have your students perform one or more of the sequence of five laboratory experiments following Exp. 24B, "Preparation and Properties of Hydroxides," in *Experiments in Chemistry*.

Section 14.1. Exhibit: the variety of fruits and foods in Sec. 14.1. Test each for acidity by indicator and taste test (*Warning to students*). Exhibit large empty bottles of industrial acids with labels. Redemonstrate acid conductivity and demonstrate properties 4, 6, and 7 (Sec. 14.5) to provide an operational definition of an acid. Discuss detail in Sec. 14.1.

Section 14.2. Discuss: importance, uses, and composition of the industrial acids. Use Table 14-1. Demonstrate preparation of dilute acids according to directions in Sec. 14.2.

Section 14.3. Demonstrations: Write simple equations for Arrhenius ionization, i.e. $HCl \rightarrow H^+ + Cl^-$. Use the electron-dot formula and a model of HCl to illustrate covalent structure. If necessary, redemonstrate the conductivity and litmus tests for glacial and diluted acetic acid; and the similar tests for HCl-benzene, HCl-benzene-water described in Chap. 13. Review the polar nature of water and write the equations for hydronium ion production in Sec. 14.3. You may wish to show that HSO_4^- is an acid (ion) by litmus test of solution of $NaHSO_4$. Retest or discuss conductivity of acetic and carbonic acids identifying these as weak acids and defining weak and strong acids.

Section 14.4. Demonstrations: Cite evidence for the existence of hydronium ion (hydrated crystals of perchloric acid. Exhibit a model of H_3O^+ and $H_9O_4^+$ (Fig. 14-1). Spatial models may be constructed from plastic spheres. Use the HRW transparencies CH-7147 (Fig. 14-2) and CH-7144 (Fig. 13-7) along with the electron-dot formulas and equations for $HCl + H_2O$. Simulate proton donation with models. Circle the electron pair representing the coordinate covalent bond in H_3O^+. Repeat for $NH_3 + H_2O$ and $NH_3 + HCl$. Use HRW transparency CH-7148 (Fig. 14-4). Reproduce Fig. 14-3 on the chalkboard. Demonstrate conductivity and litmus tests for NH_3 -benzol, NH_3 = benzol-water. Hold bottles of NH_3 and HCl mouth to mouth. A cloud of HCl forms toluene (toluol) or benzol solutions of HCl and NH_3 are mixed with the evident formation of a cloudy suspension. Permit the cloudy material to settle (this takes time in a stoppered bottle), and observe the deposited solid NH_4Cl. Develop the Brønsted definition of an acid on the basis of these activities. In the reaction of $NH_3 + H_2O$ identify H_2O as the acid. Utilize the concept of hydrogen bonding as in transparency, CH-7144.

Section 14.5. Demonstrate: the properties of aqueous acids that have not been previously performed, i.e. 5, 7, 8. Develop the ionic and net equations in Sec. 14.5 including those for the successive ionizations of polyprotic acids.

Section 14.6. Activities: List representative selection of formulas of binary acids (Table 14-2) and oxyacids (Table 14-3). Have students recognize that binary acids contain only two elements. You may wish to use the term, ternary with the oxyacids since they contain three elements, one of which is oxygen. Use the chalkboard for reproduction of Fig. 14-6, Electron-dot formulas of the oxyacids of chlorine. Arrange the oxidation states, formulas and names of the 5 chlorine acids at the left of a table. At the right, have students write the formulas of the sodium compounds of each acid that could be formed by reaction with (a) sodium, (b) sodium hydroxide. Later the names of these salts are filled in (Sec. 14.15). Drill the formula and name of the acids listed at the bottom of p. 268. Emphasize the formulas of the *ous* and *ic* acid pair in Table 14.3.

Section 14.7. Laboratory Exp. 24A: "Preparation and Properties of Acids." Use student results and redemonstrate if required. Define anhydride on this data. Write equations. Evaluate the anhydride method of preparation as applicable only to oxyacids, while the salt + sulfuric acid method may be used for both classes of acids, i.e. H_2SO_3 from SO_2 and from Na_2SO_3. Demonstrate the formation of a binary acid by direct union, burning H_2 in Cl_2. Use moist blue litmus, first in the Cl_2, then in the product.

Section 14.8. Exhibit: the common household

bases. Have students read the labels and the antidotes for household ammonia and lye. Test these with litmus as well as KOH and $Ba(OH)_2$ solutions. State the Arrhenius definition of a base. Reverse or invert HRW transparency CH-7145 (Fig. 13-11) to show that the modern version of neutralization is the reaction between hydronium and hydroxide ions. After establishing the hydroxide ion as the most common basic ion in aqueus solutions, develop the Brønsted definition of a base. Use the general format for a protolytic reaction for Sec. 14.8 reactions:

acid + base → acid + base or its useful equivalent
$A_1 + B_2 \rightarrow A_2 + B_1$.

Using this format identifies the various species as Brønsted acids or bases.

Section 14.9. Demonstrate: conductivities for various hydroxides to compare strengths. Use indicators, phenolphthalein and methyl orange. Have students detect slippery feeling of soap which is tested with litmus. Use $CO_2 + Ca(OH)_2$ as demonstration reaction for Sec. 14.9-6. For property 7, prepare and test $Al(OH)_3$ as described in Laboratory Exp. 24B, "Preparation and Properties of Hydroxides." Develop the equations for its amphiprotic behavior. $Zn(OH)_2$ may be treated in the same way. Produce a precipitate of $Fe(OH)_3$ from either NaOH or $NH_3(aq)$ and $FeCl_3$. Demonstrate that it is not amphiprotic.

Section 14.10. Laboratory Exp. 24B: Demonstrate the slaking of quicklime and test $Ca(OH)_2$ with litmus. Identify and define basic anhydrides. You may wish to use this form of the equation in Sec. 14.10:

$O^{--} + H_2O \rightarrow OH^- + OH^-$
base + acid → acid + base

Refer: To Toon, *Foundations of Chemistry*, Table 15-1 for placement of O^{--} (O^{2-}) ion in the table of Relative Strengths of Acids and Bases.

Section 14.11. Discuss: the O—H bond, its ease of breaking amphoterism, and acidic behavior as it relates to the oxidation states, and the high electronegativity of the central atom. Write the two equations for $HClO_3$ in Sec. 14.11. Give examples and draw conclusions about behavior of ionic oxides and molecular oxides (hydroxyl groups) see Sec. 14.7 and use chalkboard equivalent of Fig. 14.9 and use as developed in Sec. 14.11.

Section 14.12. Demonstrations: Develop the concept and define conjugate acid, conjugate base, and conjugate acid-base pair by writing equations and organizing or identifying by the scheme:

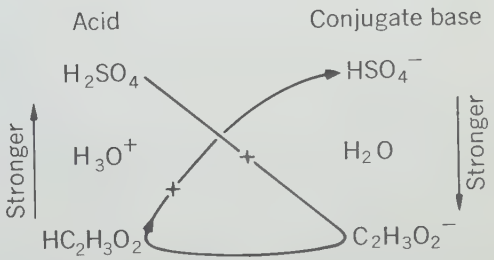
$A_1 + B_2 \rightarrow A_2 + B_1$.

For selected acids, develop part of Table 14-5 using formulas to complement the equations, i.e.,

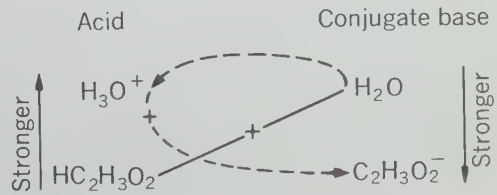
Equation

A_1	B_2		A_2	B_1
HCl	+ H_2O	→	H_3O^+	+ Cl^-
Acid				Conjugate base

Conductivity tests have established that strong acids are completely ionized and weak acids incompletely ionized. In the developing table, place the strong acids above H_3O^+ and the weaker acids below. Laboratory Exp. 27, "Relative Strength of Acids and Bases." Indicate that experimental evidence—conductivity, indicator reactions, and behavior in solvents other than water has been used to establish the relative positions of the acids (and conjugate bases). Mention the leveling effect of water on strong acids (refer to Vander Werf, *Acids, Bases, and the Chemistry of the Covalent Bond*, pp. 22-25 for details and representative diagrams). Demonstrate: that a strong acid like H_2SO_4 reacts practically completely with ions like $C_2H_3O_2^-$, CO_3^{--}, HCO_3^-, HS^{--}, to yield the weaker acids and the weaker base HSO_4^- or SO_4^{--}. This serves to establish the statement that protolytic reactions favor the production of the weaker acid and the weaker base. Transparencies and chart: C T 600-67, "Relative Strengths of Acid-Base Conjugate Pairs"; R 21759, "Chart of Acid and Base Strengths"; or the Welch chart, 4839, "Chart of Relative Strengths of Acids and Bases." Apply the text statement to the hydronium ion-hydroxide ion neutralization. Reference and demonstrations: Below are two schematics which may be used with Table 14-5. The first schematic represents: stronger acid + stronger base which react almost completely to form the weaker acid and weaker base.



Reactions between a weaker acid and a weaker base react to a slight extent to form the stronger species.



Note: the dashes indicate that the reaction proceeds to only a slight extent, i.e. acetic acid is a slightly ionized, a weak acid.

The Welch wall chart 4839, Chart of Relative Strengths of Acids and Bases may be used instead of the transparencies. Have students explain why the neutralization of a strong acid and a strong base is practically complete. You may wish to test a solution of NaCl (established as a product of strong acid-strong base neutralization with litmus. Then test a solution of sodium acetate or sodium carbonate with red litmus and challenge the students to explain the result. The second schematic deals with reactions between weaker acid and weaker base, which reactions proceed to only a slight extent to form the stronger species. The example deals with the ionization of acetic acid which has been shown to be a weak acid, i.e. the ionization takes place to a very limited extent. Test a solution of sodium acetate or sodium carbonate with litmus and challenge the students to explain the results.

Section 14.13. Review: the series of equations in Sec. 14.13 representing neutralization and leading to the definition of a salt. Demonstrations: You may wish to demonstrate neutralization by mixing 50 ml of solutions of NaOH and HCl of equal molarity and evaporate to dryness. See Alyea, pp. 146, 201, 212 for methods for demonstrating production of NaCl by direct union. Refer to Table 4-6 for precipitates (salt solubility generalizations), Tables 11 and 12 Appendix. Transparency: C T, 600-44 "Semi-Quantitative Solubility of Compounds."

Section 14.15. Activities: To establish rules for naming salts of the oxyacids, complete the table of the acids (and salts) of chlorine. Have students name the salts they recognize, i.e. they may be able to name $NaClO_3$ from previous experience with $KClO_3$. Have them suggest a name for $NaClO_4$, since its acid name also ends in *ic*. Have a student read the label on a bottle of a bleach which mentions sodium hypochlorite. This is used in changing the *ous* ending of acids to the *ite* ending for the salts. Refer to Table 14-7 for Stock system and use Table 14-8. Films and filmstrips: As a summary CORF, *Properties of Acids, Bases, and Salts;* or the filmstrip DG 636, "Acid Base Theories."

Films:

MLA CHEM 4130 *Acid-Base Indicators* (16 mm, 19 min, 8, c, s)
CORF *Properties of Acids, Bases, and Salts* (16 mm, 28 min, b/w, c, s)
CORF *Acids, Bases, and Salts*, (16 mm, 22 min, b/w, c, s)
INDU *Ionization and Ionic Equilibrium* (16 mm, 15 min, b/w, c, s)
AIM YF-236 *Ion Removal by Metathesis* (16 mm, 4 min, c, s)
YF-237 *Solubility Product* (16 mm, 7 min, c, s)
LBF *Titration Curves* (4 min, 8, c, silent)

Loops:

HR 04-96372 *Buffer Solutions* (s-8, 4 min, c, s)
HR 04-96356 *pH Meter* (s-8, 4 min, c, s)
W *The pH Meter* (s-8, (T or K), 4 min, c, s)
W *Titration* (s-8 (T or K), 4 min, c, s)
HRW *Ionic Reactions in Aqueous Solutions* (10 loops, s-8, c, s)
MLA CHEM STUDY Loop: *Acid-Base Indicators* (s-8, 4 min, c, s)

Transparencies:

HRW CH-7144 Ionization of Polar Hydrogen Chloride
HRW CH-7147 Dissolving Hydrogen Chloride in Ammonia
HRW CH-7145 Formation of Hydrogen Bond between Two Water Dipoles
HRW CH-7146 The Apparent Number of Ions Present May Be Fewer than the Actual Number
HRW CH-7148 Dissolving Ammonia in Water
HRW CH-7149 Relationship between pH and Hydronium-ion Concentration
HRW CH-7150 Acid-Base Titration Curves
K & E 30-6120-4 Hydrogen Bonding in Water (1 + 4)
K & E 30-6120-14 Ionization Reactions in Water-Liquid Solution (2 + 1)
K & E 30-6120-15 Ionization Reactions in Water-Gas (2 + 2)
K & E 30-6120-57 Naming Acids and Salts (1 + 7)
K & E 30-6120-71 Hydrolysis of Dissolved Salts (1 + 6)
C C T-600-21 Naming Acids and Salts
C C T-600-22 Normal Solutions in Titrations
C C T-600-28 pH Indicators (a)
C C T-600-29 pH Relations (b)
C C T-600-44 Semi-Quantitative Solubility of Compounds
C C T-600-66 Conjugate Acid-Base Pairs
C C T-600-67 Relative Strengths of Acid-Base Conjugate Pairs
C C T-600-68 Solubility Product Principle

TEACHER NOTES CHAPTER 15

Activities and Demonstrations: Although the chapter calls for a great deal of problem work, the related mathematics is relatively simple and the use of logarithms is a matter of choice. It affords ample opportunity to review the mole concept and its application to the preparation of solutions, use of equations, and development of the concepts of chemical equivalents and normal solutions. Exhibits of various mole quantities, fractional and multiple, show the facts on which the mathematics depend. The large number and variety of student experiments and audio-visual aids provide real or vicarious experience. The graphs of experimental results (titration) and various tables (indicators) provide the basis for related demonstrations and may also be used as the basis

on which students can apply the concepts in this chapter. Student experiments are: Exp. 25, "Hydronium Ion Concentration, pH;" Exp. 26, "Brønsted Acids and Bases, Indicators;" Exp. 27, "Relative Strengths of Acids and Bases;" Exp. 28, "Titration of an Acid and a Hydroxide, Heat of Neutralization;" Exp. 29, "Percentage of Acetic Acid in Vinegar." Experiments from *Scientific Experiments in Chemistry* are: Exp. 30, "Elementary Volumetric Analysis;" Exp. 31, "Neutralizing Phosphoric Acid;" Exp. 52, "Titration Using Chelation."

Section 15.1. Review: the mole concept. Review and demonstrations: molality and make up 1-*M* solutions of NaCl, $C_{12}H_{22}O_{11}$, C_2H_5OH. Measure the resultant volumes of the sucrose and ethanol solutions. Discuss the number of solute molecules per kg H_2O versus the number per liter of solution. Make up 1-*M* solutions of sucrose and ethanol using a volumetric flask. Pour out 100 ml of each of these 1-*M* solutions. Discuss the number of solute molecules in the 100 ml. Make up 1-*M* solutions of NaCl and KCl. Discuss the number of ions: K^+, Cl^-, and Na^+ in each solution. Weigh out 0.5 and 2.0 moles of NaCl and make up 0.5-*M* and 2.0-*M* solutions. Display volumetric flasks of different capacities. Use to make representative volumes of desired molarity. Demonstrate preparation of solutions of fractional molarities by dilution. Use the equation $M_1V_1 = M_2V_2$ to illustrate how either the new molarity or new volume may be calculated. Discuss the H_2SO_4 example in Sec. 15.1. Another relationship which is related is: Molarity = $\frac{\text{moles}}{\text{liters}}$ and therefore Moles = Molarity × liters. Question 1; problem 2-7, 16.

Section 15.2. Questions and problems: Activities: Develop the definitions of one equivalent of an acid and one equivalent of a base (in proton transfer reactions) using the examples for neutralization in Sec. 15.2. Use the same procedure for HNO_3 and $HC_2H_3O_2$ to develop the fact that one equivalent of a monoprotic acid is the same as one mole of the acid. Then consider the meaning of one equivalent of diprotic and triprotic acid in *complete* neutralization in terms of the number of H_3O^+ ions that one mole of the respective acids (H_2SO_4 and H_3PO_4) can supply. Repeat the activities above for the bases NaOH, KOH, $Ca(OH)_2$. Treat the topic of incomplete neutralization of di- and triprotic acids as in Sec. 15.2 to establish the concept: Equivalents (equivalent weights) are not constants, they are a function of the nature of the reaction.

Activities: For chemical equivalents of a reactant that is neither an acid nor a base: (a) *For elements:* Display 1 mole of Na, Ca, Al (or others, i.e., Mg, Zn, etc.) and for Ca 20 g, for Al 9 g). Have students write the equations for electron loss and develop the mole and mass relationships for each equation. Use these activities to lead to the determination of one equivalent of an element by dividing the mass of one mole by the *change in oxidation state* these atoms undergo in a chemical reaction. For balance, repeat with nonmetals such as Cl_2 and S. To show that equivalents may differ, depending on the particular reaction, try: $H_2 + S \rightarrow H_2S$ and $S + O_2 \rightarrow SO_2$. (b) *For salts:* As a lead-in to the consideration of salts, have the students reconsider the previous equations for Na, Ca, and Al and the number of moles of Ca^{++} and Al^{+++} ions (and their masses) that are equivalent to one mole of Na^+ ions. Starting with NaCl have the students write dissociation equations for Na_2SO_4, Na_3PO_4, $Al_2(SO_4)_3$. Using the Na^+ ion equivalent as the foundation, illustrate the calculation of one equivalent successively of NaCl, $CaCl_2$, Na_2SO_4, empirical $AlCl_3$, Na_3PO_4, $Ca_3(PO_4)_2$. These activities may be used to formulate the rule for calculating the equivalent of a salt. For some students it is more strategic just to state the rule and have the students apply it. For application, apply the rule to $Al_2(SO_4)_3$, $Ca_3(PO_4)_2$. Questions 2–5, problem 8.

Section 15.3. Questions and problems: Activities: If the determination of salt equivalents has immediately preceded the introduction of normality, use volumetric flasks to make up solutions of salts containing a variety of equivalents of the salts in one liter of solution. Measure out various volumes of these solutions: i.e., 100 ml, 250 ml, 500 ml of NaCl. Have students calculate the number of equivalents in each sample of solution, developing the relationships: normality = $\frac{\text{equiv}}{\text{liter}}$, $N = \frac{\text{equiv}}{\text{V(liters)}}$, equiv = $N \times$ V (liters). Have students calculate the mass of solute in each solution. Activities: Illustrate dilution operations for preparing smaller normalities from a 1-*N* solution of one or more salts. Develop the relationship: $N_1V_1 = N_2V_2$. Illustrate that ml may be used instead of liters if both V_1 and V_2 are in milliliters. Have students read the data (percent HCl and density) on the label of the concentrated acid (Fig. 5-2). Also refer them to Table 14-1. Have them calculate the mass of HCl in each milliliter of the concentrated acid. Develop the relationships between molarity and normality for acid solutions. i.e., 1-*M* HCl = 1-*N*, 1-*M* H_2SO_4 = 2-*N*, 1-*M* H_3PO_4 = 3-*N* and illustrate for solutions other than 1-*M*. Refer: to Table 15-1 for summary and comparison of methods of expressing concentration of solutions. Questions and problems: questions 17–20; problems 9, 10, 12, 18, 19.

Section 15.4. Demonstrate: the conductivity of water first using a 25-watt bulb (water a nonelectrolyte) and then a Ne or Ar lamp (autoprotolysis). Insert red and blue litmus paper to show the neutrality of water and the equal concentrations of H_3O^+ and OH^-. Write the equation for autoprotolysis refer to Sec. 13.8 and Sec. 14.9-7 (amphiprotism). Use the HRW transparency CH-7145 (Fig. 14-11).

Section 15.6. Demonstrations: Dependent on the mathematical background of the class, you may go through the mathematical operations and pH expressions in the first part of Sec. 15.6. The functional equa-

tion is pH = – log $[H_3O^+]$. For all types of classes go through both types of Sample Problems which involve an integral power of 10. Where appropriate, use the type of Sample Problems dealing with hydronium ion concentrations that are not integral powers of 10 and which require logarithmic calculation. Use the HRW transparency CH-7149 (Fig. 15-3). This is used in estimating pH from $[H_3O^+]$ or $[H_3O^+]$ from pH. You may wish to use pH paper or a pH meter to complement the results of calculations (and to validate the transparency CH-4149 for Fig. 15-3 relationships). **Tables and problems:** Use Table 15-3; problems 13, 22–24; and additional problems in Castka, *Chemistry Problems*, pp. 195–201. **Laboratory Experiments** 25, "Hydronium Ion Concentration, pH."

Section 15.7. **Demonstrations:** Review and apply the ion-product of water concept as in the introductory paragraphs of Sec. 15.7. Demonstrate the addition of 0.1 mole of solid NaOH to a liter of 0.1-*M* HCl or use 0.01 mole of the NaOH with 100 ml of the acid. Use phenolphthalein as the indicator. The complementary demonstration using solutions and burets may also be used, possibly as a lead-in to the framing of the definition of titration in which the HCl concentration is theoretically unknown.

Follow up with the calculation series for determination of the molarity of the NaOH solution used in titrating 10.0 ml of 0.01-*M* HCl solution and stressing the mole equivalence represented in the neutralization equation. Repeat with the 2 NaOH + H_2SO_4 neutralization and evolving the four operations required in the mathematical determination of the molarity of a solution of unknown concentration. Solve the Sample Problems for HCl and $Ba(OH)_2$. See Alyea, p. 16 (E6-14) for this demonstration. Discuss the importance of standard solutions. Develop the relationship $V_1N_1 = V_2N_2$ on the basis of the equality: equiv of acid = equiv of base, since solutions of the same normality are always chemically equivalent. Before dealing with the strong base-weak acid neutralization, demonstrate the basic character of a solution of sodium acetate (litmus, phenolphthalein, pH paper). **Laboratory Experiments:** Exp. 28, "Titration of an Acid and a Hydroxide, Heat of Neutralization;" Exp. 29, "Percentage of Acetic Acid in Vinegar;" *Scientific Experiments in Chemistry*: Exp. 30, "Elementary Volumetric Analysis," Exp. 31, "Neutralizing Phosphoric Acid."

Section 15.8. **Exhibit:** as many actual examples as possible of Fig. 15-7. **Laboratory Experiments:** Exp. 18, "Brønsted Acids and Bases, Indicators;" Exp. 27, "Relative Strengths of Acids and Bases."

Section 15.9. **Demonstrate:** the use of short-range pH paper such as Hydrion D(11–13) with solutions of $Mg(OH)_2$, $Ca(OH)_2$, $Ba(OH)_2$. Then check with a pH meter. The CHEM film 4130 determines the equilibrium constants for four indicators which are arranged in order of decreasing acid strength.

Films:
- MLA CHEM 4130 *Acid-Base Indicators* (16 mm, 19 min, c, s)
- AIM YF-236 *Ion Removal by Metathesis* (16 mm, 4 min, c, s)
- LBF *Titration Curves* (8, 4, min, c, s).

Loops:
- MLA CHEM STUDY 4002 "Acid-Base Indicators" (S-8, 4 min, c, s).
- HR 04-96356 "pH Meter" (s-8, 4 min, c, s)
- 04-96372 "Buffer Solutions" (s-8, 4 min, c, s)
- "Titration" (s-8, 4 min, c, s)
- W "The pH Meter" (s-8, T or K, 4 min, c, s)
- "Titration" (s-8, T or K, 4 min, c, s)
- H E or E 84-0082/1 "Use of Buret: Titration" (s-8, 4 min, c, s)
- S (W. B. Saunders Co) General Chemistry Laboratory Films 9, Volumetric Analysis: *Acids and Bases* (s, c)

Filmstrip:
- DG "Acid-Base Theories" (40 f, c)

Transparencies:
- HRW CH-7149 "Relationship between pH and Hydronium-ion concentration" (Fig. 15-3).
- HRW CH-7150 "Acid-Base Titration Curves" (Fig. 15-4).
- HRW CH-7145 "Formation of Hydrogen Bond between Two Water Dipoles."
- K & E 30-6120 "Hydrogen Bonding in Water" (1 + 4)
- C T-600-22 "Normal Solutions in Titrations"
- T-600-28 "pH Indicators (a)"
- T-600-29 "pH Relations (b)"
- McW Unit VIII: "Neutralization."

Programmed instruction:
- Barrow, G. M., et al., *Understanding Chemistry*, *I Chemical Quantities*, Benjamin, 1967, pp. 95–121; *III Chemical Reactions*, Benjamin, 1966, pp. 1–40.
- Kass, G. A., *Individual Study Program in Chemistry*, *Volume IV*.
- *Acid-Base Chemistry*, Benjamin, 1969, pp. 40–96, 129–158.

TEACHER NOTES CHAPTER 16

Chapter Note: **Activities and Demonstrations:** The experiments for this chapter are Exp. 30, "Adsorption;" Exp. 31, "Carbon;" and Exp. 32, "Carbon Dioxide." These experiments should be scheduled to permit the student to gain first-hand experience before the textbook lessons. You may wish to demonstrate parts of the experiments which were either omitted in the laboratory or require further consideration.

Section 16.1. **Laboratory Experiments:** 30, "Adsorption;" 31, "Carbon;" 32, "Carbon dioxide." **Exhibit:** forms of carbon, coal, petroleum, common organic chemicals, medicinals and synthetics. These are representative of the abundance of carbon compounds

and lead to the definition of organic chemistry.

Reference: *Chemistry*, Reprint 55, "Coal." Exhibit: Periodic table and locate C. Review electron configuration, and electron-dot symbols for C in ground state. Exhibit models or samples of carbon in bonded state: tetrahedron, CH_4, CCl_4, $CHCl_3$, CO_2; tetravalency; and the symbol $\cdot\dot{C}\cdot$ Review schematic (Chap. 6 TE) for sp^3 hybridization.

Exhibit: ball-stick models of diamond, graphite, simple hydrocarbons, benzene, crystal models: diamond, graphite; samples of forms of carbon; allotrops of S, P. Demonstrations: formation of S, i.e., rhombic, monoclinic, amorphous; red to white P (Alyea p. 177); ozone (Alyea pp. 57, 58, 127). Definition: allotropy.

Section 16.2. Exhibit: diamond model, tetrahedron, phonograph needle. Electron-dot diagram and network structure of diamond (Fig. 16-2, HRW transparency CH 7154). Demonstration: diamond scratches glass, MOH scale of hardness.

Sections 16.5-16.6. Exhibit: samples of graphite-chunk, powdered, lubricant, model (Fig. 16.4). Demonstrations: Make an erasable mark on paper with graphite-chunk, erase; duplicate with pencil, properties as in Exp. 31, "Carbon." Exhibit: Graphite greases and lubricants, crucibles, electrodes and discuss properties and uses of each. Include use as moderator in atomic reactor (Chap. 30). Discuss: properties and structure-layer and weak interaction forces, $sp^2 - sp^2$ bonding, *Chemistry Problems*, p. 538, Q. 4.

Section 16.7. Exhibit: samples of forms of amorphous carbon and representative uses—charcoal chunks and brickets, dry-cell electrode, petroleum coke, coke, portions of a tire, ink, paints, carbon paper, phonograph records, etc. Review: Exp. 31, "Carbon," destructive distillation and properties. Demonstrations: destructive distillation—charcoal and coke. Demonstrations: The following sequence of demonstrations may be used to develop understanding of the process of destructive distillation. A wooden splint held in an open flame, and the same splint heated in a test tube, challenges student interest. Three or four splints broken in half may be heated strongly with the mouth of the test tube held downward over a watch crystal. As the vapor is liberated it is ignited. Blue litmus paper is used to indicate the acid nature of the distillate. The resultant charcoal strips are removed from the test tube and the manner of burning compared with the manner of a wooden splint. The porosity of the charcoal may be used to explain the steady glow of the charcoal splint. The definition of destructive distillation is developed with emphasis on decomposition by heat in the absence of air, followed by the condensation and collection of volatile materials. More complex setups may also be used, such as collecting the condensate in a bottle and burning the volatile material at another outlet from the bottle.

The production of coke by destructive distillation may also be demonstrated. If the soft coal is powdered and heated with a Fisher burner, good results are obtained after a brief heating. The coal tar collects in the bottle and the coal gas may be ignited at the bottle outlet. A piece of moist red litmus paper suspended from the cork inside the bottle turns blue, indicating the production of ammonia. The ammonia is generally liberated toward the end of the process rather than at the beginning. Some teachers dramatize this by strongly heating the open test tube and permitting the vapor to reach a foot-square piece of moistened red litmus paper, which turns blue. The properties of coke as a fuel and as a reducing agent should be compared with charcoal. The reduction of CuO may be redemonstrated if desired and the production of CO_2 shown by the use of $Ca(OH)_2$. The three main forms produced by destructive distillation should then be briefly reviewed. Soot may be produced by deposit from a candle flame and carbon black by deposit from a yellow bunsen flame. The importance of the latter in automobile tires should be emphasized. Demonstration: carbon black production—hold a test tube in the yellow burner flame and note carbon deposit.

Section 16.11. Demonstration: Collect test tube of dry NH_3 by displacement of Hg. Introduce pellets of activated charcoal through the Hg into the gas. Add pellets of activated charcoal to bottle of $Br_2(g)$, shake and observe results. Remove pellets of charcoal and heat in test tube. See Experiment 30 Instructor Experiment for other adsorption demonstrations. The student section of Experiment 30 deals with the adsorption of acetic acid. In Experiment 31, "Carbon," students perform an experiment representative of sugar refining which may be repeated as a demonstration. Test the relatively colorless filtrate with Fehling's or Benedict's solution.

Section 16.17. Demonstrations: Expose limewater in a watch glass on the table top and have students observe the formation of the crust of $CaCO_3$ on the upper surface. Burn a splint in a bottle of air and test result with limewater. Have students blow breath through a straw into limewater in a test tube.

Section 16.13.-16.16. Demonstrations: The topic of carbon dioxide provides an excellent opportunity for the use of a series of fire-extinguishing demonstrations (Sec. 16.18) which may well be used before dealing with its preparations and properties. Set up the Instructor Experiment for Exp. 32, "Carbon Dioxide," during the laboratory period so that the resultant production of CO_2 by fermentation may be presented in the class period. Carbonates which may be heated to yield CO_2 are $MnCO_3$, $CuCO_3$ or $CuCO_3 \cdot Cu(OH)_2$. The easiest demonstration is the thermal decomposition of $NaHCO_3$. Chunk limestone may be heated on the grid of a high temperature burner. After cooling, H_2O is added to the CaO produced and the litmus test for $Ca(OH)_2$ performed. The laboratory process may be

generalized by using a variety of carbonates or hydrogen carbonates, and acids. Add HCl to marble, seashells, and limestone for observable gas production. Ask students how they would produce CO_2 by using materials in the home—vinegar or lemon juice with baking soda, adding water to baking powder, fermentation using dry yeast. Develop board equations representative of various production methods.

Section 16.15. Exhibit: a ball-stick model of CO_2 and have students explain the nonpolarity of the molecule. Develop the electron-dot formula(s) and the electronic formulas for the resonance hybrids.

Section 16.16. Demonstrations of physical properties: Use a Kipp generator as source of CO_2 for large-scale demonstrations. Variations of Procedure 4 in Exp. 32, "Carbon Dioxide" include: Set up of a very large battery jar with a staircase arrangement of burning candles. Extinguish them in succession by pouring CO_2 (previously collected by displacement of air in a large cylinder). It fills the jar from bottom to top. Exhibit the contents of a disassembled soda-acid extinguisher. Use a laboratory model to extinguish a newspaper fire in the sink. Demonstrate that a soda-acid extinguisher does extinguish an oil (benzol or Benzene—*Caution*). fire. Use a small-scale foam extinguisher to extinguish an oil fire, directing the foam nozzle against the side of the vessel rather than directly at the burning oil. A benzol fire in a large tray is dramatically extinguished by using the liquid CO_2 extinguisher. The production of Dry Ice may be shown by permitting the CO_2 from the liquid CO_2 extinguisher to fill a paper bag. Shake out the contents of the inverted bag. Relate the laboratory preparation of CO_2 to the soda-acid and foam extinguisher reactions (Sec. 16.18).

Section 16.17. Demonstrations of chemical properties: Plunge a burning splint into a bottle of CO_2. Follow up by inserting a burning Mg ribbon (use tongs and Caution) into another bottle of CO_2. Have students note both the black and white deposits on the bottle and explain the burning. Heat a sample of soda water. Test with litmus before and after heating. Pass CO_2 into limewater. Do this until a precipitate forms and redissolves. Develop formula and ionic equations in Sec. 16.17.

Section 16.18. Discuss: other uses besides firefighting as in Sec. 16.18. Discuss photosynthesis and the equation for it.

Sections 16.19 and 16.24. Discussion: introduce carbon monoxide by the statement that both carbon dioxide and carbon monoxide may produce asphyxiation and having students explain the differences in their physiological actions.

Section 16.20. Demonstrations: Water gas preparation (Alyea, pp. 9, 97). Refer to burning of gaseous product of destructive distillation (wood and soft coal). Use a diagram of a coal furnace or an old-fashioned potbelly stove showing flames burning at the top of the coal layer. Develop a series of equations and place on the diagram, (1) burning of coal on bottom grate (2) reduction of CO_2 by C as it passes through the hot coal, and (3) burning the CO at the top of the coal layer. The first two equations may be combined into a single equation which represents the incomplete combustion of coal or coke. Closing the air entrance at the side of the stove leaves CO representative of the manufacture of producer gas. Alternate closing of the draft to the chimney shows how a furnace may produce CO which could leak into the home. Have students write equations representing incomplete combustion of octane. If the laboratory preparation is demonstrated, emphasize the precautions in using hot H_2SO_4 and collecting the poisonous and combustible CO. See Alyea, p. 36, 16-9.

Section 16.21. Develop: chalkboard structures for the resonance hybrid of CO and discuss details in Sec. 16.21.

Sections 16.22-16.23. List: The physical and chemical properties using equations for the latter.

Section 16.24. Exhibit: A gas mask and discuss failure of activated charcoal to adsorb CO and use of other chemicals needed.

Films:

EBF 1954 *Carbon Fourteen* (12 min, s, b/w)
CORF *Carbon and Its Compounds* (10 min, s, b/w, c)
DUART MIS 676 *Control of Air Pollution* (16 mm, 5 min, c, s)
MIS 677 *Sources of Air Pollution* (16 mm, 5 min, c, s)
MIS 678 *Effects of Air Pollution* (16 mm, 5 min, c, s)
SUTH *The Modern Chemist-Diamond Synthesis* (16 mm, 13^+ min, c, s)
CHUR *The Photosynthesis and Respiration Cycle* (16 mm, 13^+ min, c, s)
GE *Manmade Diamond* (16 mm, 11 min, c, s)
YF 212 *Fire Safety in the Laboratory* (16 mm, 2 min, b/w, s)

Filmstrips:

DG 650 "The Chemistry of Carbon and Silicon" (40f, c)

Loops:

Photosynthetic Fixation of Carbon Dioxide I & II (s-8, 4 min, c, s; 8, 6 min, c, s)
E 81-5118/1; 81-5126/1 or
ML 6106, 6107

Transparencies:

HM 2-10529 "Hybrid Orbitals"
2-10535 "The Methane Molecule, CH_4"
2-10537 "Carbon Dioxide Structure"
K & E 6120-12 "Allotropes of Oxygen"
K & E 6120-13 "Allotropes of Sulfur"
K & E 6120-16 "Hybridization"
K & E 6120-66 "Electric Resistance Furnace"
C T-600-57 "Resonance Bonding" (Electron Dot Form)

C T-600-58 "Resonance Bonding" (Dash Bond Form)
C T-600-59 "Organic Chemistry—An Elementary Survey."

Slides:
W "The Properties of Carbon"
W "The Oxide of Carbon"
W "The Fire Extinguisher"
W "Coal"
W "Chain Hydrocarbons"
W "The Allotropes of Phosphorus"
W "Preparation and Properties of Ozone"

TEACHER NOTES CHAPTER 17

Chapter Note: Activities and Demonstrations: Many of the generalizations in this chapter such as reasons for the abundance of carbon compounds, homologous series, general formulas, isomers, and properties of organic *vs* inorganic compounds, are developed best by the use of samples, transparencies, models, formulas, tables such as Table 17-1, and demonstrations (or student experiment results). The inductive development is accompanied by the step-by-step accumulation of data on the blackboard, which students record in notebooks. Separate recording of generalizations are written on the board and recorded by students as they emerge from discussion. Some examples are given below. The gradual presentation and accumulation of facts leading to pertinent generalizations help students' mastery of organic chemistry. You may wish to demonstrate the "Reversion of the Synthesis of Urea," Alyea, p. 152, 209 in conjunction with the vital-force theory. In this demonstration, a solution of urea is tested for electrical conductivity and precipitation with $AgNO_3$ solution. Another solution of urea is boiled for several minutes and the conductivity and precipitation tests are performed. The boiling produces ammonium and cyanate ions (NH_4^+, CNO^-). Another reagent which may be used is iron (III) chloride which produces a red color with the solution which has been boiled. Try cooling the boiled solution in ice and retesting. Details of the Wöhler synthesis are described in *Principles and Practice of Organic Chemistry*, Lucas, H. W., and Pressman, D., Wile, 1949, p. 310. An interested student may wish to attempt this experiment as a project.

Section 17.1. **Define:** Organic chemistry (Sec. 16.1). **Transparency:** C T 6000-59, "Organic Chemistry —An Elementary Survey." Have students propose a question about reasons for the existence of the great number (3,000,000 + 100,000 yearly) of carbon compounds. To show that carbon is not among the 10 most abundant elements, refer students to data in Chap. 2. Review details of tetravalency of carbon, covalent bonding, diamond structure, and hybridization developed in Chap. 16. One possible chalkboard listing of reasons for abundance of carbon compounds is: 1. Tetravalency of carbon, hybridization. 2. Carbon atoms can combine with each other in (a) chains—straight, branched, (b) networks (diamond), (c) rings (benzene). 3. Carbon atoms combine by (a) single bonds, (b) double and triple bonds. 4. Isomers (define). 5. Combination with other elements (substitution, addition).

Section 17.3. **Activities:** Review calculation of empirical formula from data in Sec. 17.3. Construct, or have students construct, models of two different substances. Have students make the waft-odor test with samples of each substance. Write the structural formulas. You may wish to demonstrate the action of sodium on ethanol and exhibit the reagent phosphorus trichloride and ethyl chloride.

Section 17.4. **Demonstrate:** The differences with the text-suggested substances or equivalent (a) *solubility:* NaCl, NH_4Cl, KNO_3 in water; oil and naphthalene + water; selected solutes in benzene; ethane and sugar + H_2O. (b) *decomposition:* heat sugar and NaCl. (c) a precipitation contrasted with fermentation (other instances later on, esterification, saponification, polymerization). (d) *molecular structure-covalence:* recall conductivity tests for ethanol and acetic acid in water, refer to urea synthesis. Mention Cautions in Sec. 17.4, refer to CCl_4, glue-sniffing.

Section 17.5. **Develop:** The concept of homologous series and general formula, for the alkanes, and reinforce with later series. **Exhibit:** paraffin, Bernzomatic type torch (Sears Craftsman model has a label—Propane Fuel), butane-fueled cigarette lighter. Light the laboratory burner and mention natural gas-methane. Now or later exhibit a set of hydrocarbon products from petroleum (some oil companies furnish such a set). **Develop:** the details of the outline table.

Name of Series
General formula

Names	*Molecular Formula*	*Electron-dot Formula*	*Structural Formula*	*Phase*

Section 17.10. **Exhibit:** Display petroleum products. A chart presently available is *Facts about Oil.* Refer to Experiment 33, "Hydrocarbons," for preparation and properties of methane. For substitution, use white mineral oil or paraffin with Br_2 in perchloroethylene (instead of CCl_3). Ligroin (a mixture of C_6 and C_7 compounds) or *n*-heptane (both from Arthur H. Thomas Co.) may be tested with bromine water. Try bringing an ultraviolet lamp across the mouth of the bottle and then blow breath across the mouth of the bottle to detect fuming HBr. (Also test with moist blue litmus paper at bottle mouth.) Addition of $KMnO_4$ solution with an alkane produces no perceptible change.

Section 17.11. **Develop:** a chalkboard table for ethene and propene similar to that for alkanes and establish general formula. **Exhibit and activities:** polyethylene and propylene laboratory beakers etc. Write the molec-

ular formulas for ethane and ethene and challenge students to write a structural formula for C_2H_4. Ask why both compound names start with *eth*. Construct ball-stick models for ethene and propene. Develop concept of sp^2 hybridization, relate to HRW transparency CH-7180, identify sp^2-*s*(C-H) bonds, sp^2-sp^2 (C-C) bonds —both of these are sigma bonds, *p-p* overlap (*pi*) bond, and relate to planar structure of ethene and 120° bond angles. **Demonstrations:** Preparation of ethene and testing with Br_2 and $KMnO_4$ solutions (Fowles, Exps. 138a, 142, 143; Alyea, pp. 47, 174). Cracking: pretest white mineral oil with Br_2 solution. Heat a sample of the mineral oil in a test tube with steel wool and repeat Br_2 test. On a larger scale, the vapor from boiling mineral oil is passed over steel wool heated in a connected glass tube. Collect the gases produced by displacement of water. Amylene or 1-pentene may also be used to demonstrate addition. Demonstrate combustibility. Write equations for preparations and reactions. Use models to illustrate addition. Use HRW transparency CH-7180 or K & E 6120-10, "Isomers of dibromoethane." Demonstrate hydrogenation and polymerization with ball-stick models.

Section 17.14. **Develop:** a table for ethyne and propyne similar to the previous tables. See Procedure 3, Experiment 33, "Hydrocarbons," for preparation and properties of ethyne. **Demonstrations:** Drop 1 or 2 chunks of CaC_2 into a test tube half-filled with H_2O. Ignite the gas at the test tube mouth. Test the remaining contents of the test tube with litmus or phenolphthalein. Contrast the flame with that of Fig. 17-10. By holding a test tube in the type flame of burner in Fig. 17-10, a deposit of carbon is produced on it.

Section 17.18. **Demonstrations:** Develop a table similar to preceding ones for benzene, toluene and xylene. Demonstrate destructive distillation of bituminous coal using a test tube, collecting bottle, jet assembly and igniting the coal gas at the jet. Use a high temperature burner to speed decomposition. Coal tar collects in the bottle. **Demonstrations:** Ignite styrofoam or articles of polystyrene which burn readily with a characteristic odor. Solid polystryene objects give a metallic ring when dropped. Use the reaction equations for Sec. 17.19 and Sec. 17.20 to the extent required by your course of study. Use Fig. 17-12, show equivalent spatial models (commercial kits), and write the hexagonal representations for naphthalene and anthracene.

Section 17.21. **Exhibit:** The wall chart (American Petroleum Institute). Distillation or Fig. 17-13 and Fig. 17-14. Discuss structure and operation of pipe still and fractionating tower. Contrast fractional distillation (physical change) with destructive distillation (chemical change). Use Table 17-2 as a summary. Demonstrate fractional distillation of a mixture of alcohol and water. See Alyea, p. 105 for demonstration of the distillation of petroleum.

Section 17.23.-17.24. **Discussion and references:**

I don't care for catalytic crackers. I prefer Premium Saltines.

Discuss the automobile-related problems of pollution, control devices in automobiles, and effects on the type of gasoline, its composition, and the costs. Have students consult the Index of recent December issues of *Chemistry* for related articles. An article in the Sept. 1972 issue of *Reader's Digest*, "The Clean-Auto Law Is the Wrong Answer," presents a readable comprehensive summary of the overall problems. Reprints are available. Discuss the octane rating of gasoline, the antiknock properties of various components (hydrocarbon classes) and the use of tetraethyl lead as in Sec. 17.23.

Section 17.25. **Exhibit:** Rubber latex, unvulcanized (if available) and vulcanized rubber, sulfur, carbon black, tires, and rubber objects. **Demonstration:** Add acetic acid to rubber latex to produce coagulation. Squeeze into a ball and bounce. A salt solution will also produce coagulation. See Alyea, p. 101 Vulcanization with Sulfur; p. 114 Electrodeposition of Rubber; p. 157. Coagulation of Rubber Latex. Construct a ball-stick model of isoprene. From *Scientific Experiments in Chemistry*, Experiment 33, "Rubber."

Section 17.26. **Exhibit:** Gloves and tubing made of Neoprene. Use models and equations for the reactions involved in its manufacture. Compare the structure of the Neoprene unit with that of the C_5H_8 unit of natural rubber. Discuss details of manufacture and unit structure of SBR rubber as in Sec. 17.28. Exhibit tires. **Demonstration:** Production of Thiokol rubber,

Alyea p. 110 (22-48s), (22-52s); Experiment 20-3, "Synthetic Rubber," *Laboratory Experiments for Foundations of Chemistry* is based on Alyea, 22-52s for production and coagulation of Thiokol latex. The Thiokol is a condensation polymer which splits off NaCl when $C_2H_4Cl_2$ reacts with sodium polysulfide. The sulfur atoms provide linkage and the production illustrates the role sulfur plays in vulcanization of natural rubber.

Transparencies: HRW Series:
- CH-7152 The four covalent bonds of a carbon atom
- CH-7153 The Methane Molecule
- CH-7154 The crystal structure of Diamond
- CH-7155 The crystal structure of Graphite
- CH-7156 Models of Methane, Ethane and Propane (Fig. 17-1)
- CH-7157 Bonding in Ethane (Fig. 17-3)
- CH-7158 Methane and Bromine undergo a substitution reaction (Fig. 17-6)
- CH-7259 Bonding in Ethene (Fig. 17-7)
- CH-7160 Ethene and Bromine undergo an addition reaction (Fig. 17-8)
- CH-7161 Bonding in Ethyne (Fig. 17-9)
- CH-7162 Bonding in Benzene (Fig. 17-11)
- K & E 6120-9 Isomers of Butane (3 + 3)
- K & E 6120-10 Isomers of Dibromoethane (3 + 4)
- K & E 6120-11 Isomers of C_2H_6O (3 + 4)
- K & E 6120-22 Covalent Bonding—Ethane, Acetylene, Ethylene
- HM 2-10535 The Methane Molecule, CH_4
- HM 2-10538 Structure of C_2H_4
- C T-600-59 Organic Chemistry—An Elementary Survey
- W (slides) (Chain Hydrocarbons)

Filmstrips:
- DG 650 The Chemistry of Carbon and Silicon (40f, c)
- DG 603 Organic Chemistry (40f, c)

Films:
- CORF *Carbon and Its Compounds* (16 mm, 11 min, b/w, c, s)
- *Hydrocarbons and Their Structure* (16 mm, 13 min, b/w, c, s)
- Shell Oil Co., *Crude Oil Distillation*

TEACHER NOTES CHAPTER 18

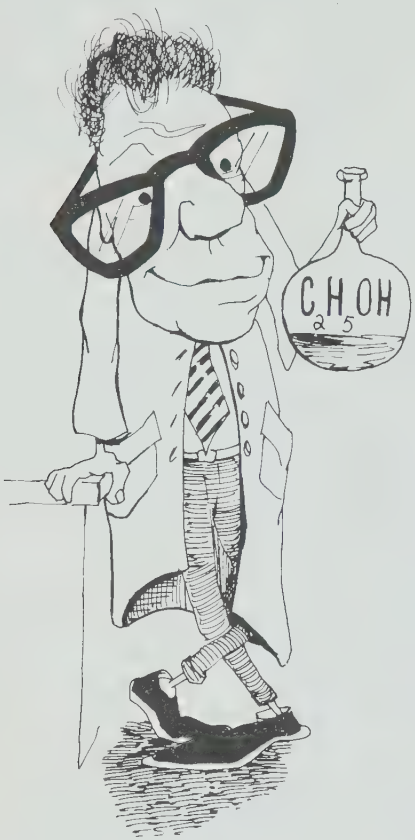
Chapter Note: Activities and Demonstrations: General formulas and characteristic distinguishing (functional) groups can be arrived at by the use of exhibited samples, formulas, and models for several of each type of compound. Discussion of pesticides and detergents as they relate to ecology may be based on student reports from articles in such magazines as *Chemistry*. Laboratory experiments from *Scientific Experiments in Chemistry*: Experiment 37, "Preparation of Esters"; Experiment 38, "Soap"; and Experiment 35, "Synthesis of a Dye." Some review of the molecular, electron-dot, and structural formulas of representative alkanes, alkenes, and alkynes will probably be required, particularly of the first four alkanes. This provides the basis for most of the chapter and understanding of the relationship of the various types of compounds in this chapter. Exhibit samples of the halogen substitution products including those with which the student may have had environmental contact. These may include Teflon and Teflon-coated pots, ethyl chloride, aerosol dispensers and iodoform.

Sections 18.1-18.3. Laboratory Experiment 34: "Hydrocarbon Substitution Products," provides for the construction of ball-stick models of representative compounds. In the classroom, many of these models may be manipulated to illustrate the preparation and properties of the compounds. The use of these models complements exhibited samples, actual experiments or demonstrations, writing molecular and structural formulas, and equations. Useful models are Ealing demonstration size sets: Molecular Orbital Models (67-7757, 67-7765) and Space-Filling Molecular Models (67-7559). Write the molecular and structural formula equations for the progressive chlorination of CH_4. List the names: monochloromethane (methyl chloride), dichloromethane, trichloromethane (chloroform), tetrachloromethane (carbon tetrachloride). Review demonstrations for substitution (Sec. 17.10 and addition (Sec. 17.13, Sec. 17.16) and use related transparencies. Allow octane or petroleum ether to react with bromine water. Illuminate the reaction with a sun lamp or 150-watt projection lamp for evidence of substitution. Write general equations using X_2 for halogen and develop general formula for alkyl halides (RX). Write the equation for alcohol + hydrogen halide. Demonstration: Prepare iodoform by shaking a dilute I_2 (aq) with ethanol (Alyea, p. 106). Perform the Beilstein test in Experiment 34, Further Experimentation. Test CCl_4 or $CHCl_3$ for chloride ion. Heat a very small sample of soda lime, add 2-3 drops CCl_4 (*Caution*). Reheat, add CCl_4 again. After cooling, add 10 ml HNO_3 (6-*M*) and test for chloride ion again.

Section 18.2. General Equations: Write equations for RX with Mg to form Grignard reagent (RMgX). Write equation for RX with OH^-, recalling the previous demonstration (CCl_4 + soda lime).

Section 18.3. Exhibit: CCl_4 (+ pyrene fire extinguisher, if available), $CHCl_3$, CCl_2F_2 (spray can, Teflon-coated pots). Write equations in Sec. 18.3, use models. Write the structural formula for C_2F_4 and the related monomer unit for Teflon. Refer to Fig. 18-1. Discuss uses and properties of alkyl halides, mentioning ethyl chloride (local anesthetic). Other exhibit materials of which the following are representative are: Fuller Brush Co., Spot Remover Spray (Can includes statement "Does not contain carbon tetrachloride" and that contents include perchloroethylene). Texize K2r also includes perchloroethylene. Exhibit paradichlorobenzene pellets in package container and Fuller Brush Co., Moth

Proofer can. Contents label mentions paradichlorobenzene as well as diethyl diphenyl dichloroethane.



Better living through chemistry.

Section 18.4. Exhibit: samples of methanol, ethanol, denatured alcohol, ethylene glycol (antifreeze), glycerol, isopropyl alcohol. Have students use common names, i.e., wood and grain alcohol, and glycerine. Light alcohol lamp and Sterno can. Develop a chalkboard table showing the relationship of the primary alcohol to the alkanes using structural formulas and names and leading to the concept: an alcohol is the hydroxyl substitution product of the hydrocarbon and may be represented as ROH. Complement by ball-stick models. Also use models for propanol and isopropanol and respective names as 1-propanol, 2-propanol. Use models to develop isomeric butanols and give example of naming, e.g., tert-butyl alcohol as 2-methyl-2-propanol. You may distinguish between primary, secondary, and tertiary alcohols by the number of carbon atoms bonded to the COH group or use alternate: $-CH_2OH$ (primary), $-CHOH$ (secondary), $-\overset{|}{\underset{|}{C}}-OH$ (tertiary). Attempt separation of ethanol by fractional distillation from fermentation experiment, Experiment 32, "Carbon Dioxide." Discuss other methods for preparing alcohols in Sec. 18.4 by using equations, models, structural formulas. Refer to Sec. 18.14 for production of glycerol by saponification. Discuss uses and properties of specific alcohols.

Section 18.5. Activities: Light alcohol lamp and Sterno if not used previously. Have students write equations for complete combustion of methanol and ethanol. Test various alcohols and their water solutions with litmus to show they do not ionize to yield hydroxide ion. React a small piece of Na with ethanol to show relationship to H_2O. Fowles (pp. 120–122) describes demonstrations for dehydration of ethanol using H_2SO_4 or H_3PO_4. Demonstrate formaldehyde production from methanol by oxidation with CuO, Experiment 34, Procedure 2.A.a. Write equations representative of other reactions of alcohol, Sec. 18.5. Demonstration of other properties of alcohols. Length of carbon chain (in 100 ml H_2O: *n*-butyl alcohol (7.9 g) *n*-amyl alcohol (2.7 g); effect of branching—*tert*butyl alcohol is completely miscible. Preparation of "solid" alcohol: add 60 ml of denatured alcohol to 10 ml of saturated solution of calcium acetate. Sample is ignited on asbestos sheet. Chemical tests distinguishing primary, secondary, tertiary alcohols: reaction with Na—1 butanol, 2-butanol, 2-methyl-2-propanol yields progressively fewer bubbles with production of white solids for last two; use of $KMnO_4$. *Journal of Chemical Education*, 30, 395 (1953).

Section 18.6. Demonstration: illustrate the methods of production of ethers by use of equations and models. Use the general formula ROR^1 and define as an organic oxide. Exhibit container of diethyl ether. Demonstrate its volatility by rapid evaporation on chalkboard. Discuss use of ether.

Section 18.7–18.8. Exhibit aldehydes: formaldehyde, formalin, paraformaldehyde, acetaldehyde and paraldehyde. Test for odors by waft test, e.g., acetaldehyde has a pearlike odor. Demonstration: See Teacher Note for Sec. 18.5 for production of formaldehyde from methanol. Mention commercial catalysts and discuss uses. Refer to Sec. 17.16 for production of acetaldehyde from ethyne. Demonstrations: 1. Fehling test and production of a silver mirror: Experiment 37, Procedure 2.D. 2. Oxidation of an aldehyde: dissolve 1 g $K_2Cr_2O_7$ in 5 ml of 3-*M* H_2SO_4. To this add a few drops of ethanol. Observe dichromate color change. Detect pearlike odor of acetaldehyde. Heat contents of tube gently and test vapors with moist blue litmus paper for acetic acid. Its odor is also apparent. 3. Condensation: plastic production—mix equal volumes of aniline hydrochloride and formalin in a beaker; stirring results in a stiff colored gel which eventually turns black. See Alyea, p. 48 (22-18) for dramatic production of a phenol-formaldehyde plastic. Use equations in Sec. 18.8 for aldol condensation and Grignard reagent reactions.

Section 18.9. Display: acetone and note odor. Demonstrations: 1. Careful heating of calcium or barium acetate with a luminous flame. 2. Methyl ethyl

ketone is produced by addition of sodium dichromate solution to a mixture of 6-*M* H_2SO_4 and 2-butanol in a reaction flask in an ice bath (*Caution*). The temperature is kept between 20°-40° C as recorded by an inserted thermometer. 3. Dissolve a sample of cellulose acetate (cleaned safety film) in acetone. Pour some of the solution on a glass plate to obtain a film of the acetate by evaporation (Use hood). 4. Separate Fehling solution tests with acetone and acetaldehyde, no reaction with acetone. Write the equation for production of acetone from 2-propanol and use models. You may wish to develop the idea that oxidation of a primary alcohol produces an aldehyde and then an acid while the initial oxidation of a secondary alcohol produces a ketone. Treat the reactions of ketones as desired according to Sec. 18.10. Write the general formula for ketones.

Section 18.11. Exhibits: samples of vinegar, lactic, citric, tartaric, oxalic, and other carboxylic acids. Test fruits (lemons, grapefruit, oranges) with litmus paper. On occasion, a bottle of cider vinegar contains a gelatinous mass (mother of vinegar bacteria); review details of fermentation of glucose yielding ethanol. Test with speed of neutralization before developing the equation and using models for the esterification an unrefrigerated sample of wine, standing for a long time, for acid with litmus paper (if test is negative, sip the wine); write equations for preparation of acetic and formic acids. Laboratory Experiment 29, "Percentage of Acetic Acid in Vinegar" fits here. Exhibit models of acetic and formic acids, identifying carboxyl group and writing general formula. Perform conductivity tests with glacial acetic and lactic acid. Test these with litmus. Breathe on litmus or moisten with water. Dilute acids and repeat conductivity test. Write equations for ionization of acetic and formic acids. Stir an aspirin in water and test with litmus (refer to TV commercials). Test aspirin and Bufferin tablets dissolved in water with pH paper, e.g., Hydrion paper A.

Section 18.12-18.13. Demonstrate: the production of methyl salicylate using a beaker for mixing salicylic acid and methanol. Add a few drops of H_2SO_4. Heat gently. Hold the beaker with tongs, walk around the room and make students recognize the oil of wintergreen odor. Prepare other esters as in Experiment 37 or from *Scientific Experiments in Chemistry*, Experiment 37, "Preparation of Esters." Exhibit various esters and have students note characteristic odors. Compare the rate and need for a catalyst in esterification mechanism in Sec. 18.12-13. Develop the general formula $R-\overset{\overset{O}{\|}}{C}-O-R^1$ for esters of carboxylic acids. Write other equations such as production of glyceryl trinitrate. Refer back to equations for sulfation of long-chain alcohols (Sec. 18.5).

Section 18.14. Exhibit: various long-chain acids, e.g., stearic, oleic, palmitic, and write the representative formulas. Exhibit various samples of fats and oils. Develop the formula for glyceryl tristearate and then the formula representing the general structure of a fat or oil. Exhibit Crisco solid and oil (label reads, "High in Polyunsaturates"). Discuss catalytic hydrogenation of vegetable oils. Test a light-colored cooking or salad oil with Br_2 in CCl_4 or in perchloroethylene for unsaturation evidence. Demonstration: Saponification from Experiment 37, Procedure 2.F. or from *Scientific Experiments in Chemistry*, Experiment 38, "Soap." Define saponification and write a representative equation. Demonstrate "salting out" by adding NaCl to a soap solution. Discuss soap production. Write the equation for hydrolysis of fats. Discuss the neutralization of the resultant long-chain carboxylic acids and varieties of soap.

Transparency:
Western slides, "Fat and Soap."

Films:
RIC, *Detergents Up to Date; What is Soap?*

Summary Films:
CHEM 4166, *Mechanism of an Organic Reaction.*
CHEM 4163, *Synthesis of an Organic Compound.*

Loops:
An Organic Mechanism, Ward, 168W1011(T) or 168W1031(K).

Filmstrips:
DG 603, "Organic Chemistry";
DG 656, "Functional Organic Groups."

TEACHER NOTES CHAPTER 19

Chapter Note: Activities and demonstrations: Because of the multiplicity of available demonstrations, experiments, and audio-visual aids it is impossible to utilize all of the suggestions. The matter of strategic choice depends on local course requirements and the general caliber of students of particular classes.

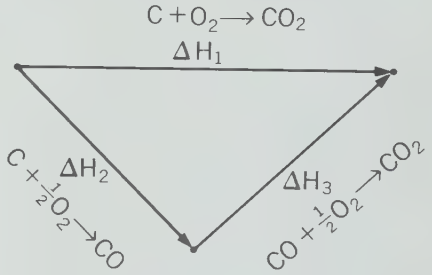
Section 19.1. Demonstrations: Introduce chemical kinetics particularly reaction rates: (*a*) *NaCl + AgNO*$_3$ (instantaneous precipitate); (*b*) *HCl to* $Na_2S_2O_3$ produces sulfur (perceptible time lapse); (*c*) *selected clock reaction:* Alyea pp. 19, 85, 130, 147, 149, 179, 220 for a sequence of steps or pathways.

Section 19.2. Demonstrations: Ask students: "What happens when hydrogen and oxygen are mixed?" They may state that an explosion occurs with the formation of water. Place labeled bottles of H_2 and O_2, previously collected, mouth to mouth. Remove cover plates. Result? Remove one bottle and ignite with flame. A mixture made by collection of 20 ml of H_2 and 10 ml of O_2 may also be made in an eudiometer and then sparked. To further illustrate the idea of mechanisms or pathways, exhibit and have students manipulate models of H_2, O_2, and H_2O. During the discussion, introduce and review concepts of bond breaking, bond making, the related energies. (See Sec.

6.9, Sec. 6.10 for bond energies). Discuss the exothermic nature of the reaction. Write the set of equations in Sec. 19.2 sequentially. Use them to define thermochemistry and heat of reaction. Define enthalpy. Emphasize reference temperature as 25° C. Discuss the representative standard state phase symbols, and the expression of heat of reaction stated first, in kcal for $2H_2(g) + O_2(g)$ and then conventionally in kcal/mole. Make sure that students understand the use of $\frac{1}{2}O_2$ and the reason for its use. Either develop the curve for Fig. 19.1A on the chalkboard or use the HRW transparency CH-7163(A). Write the related equation using the ΔH notation (mention *change* in heat content—enthalpy). Explain the negative sign for exothermic reactions. Refer students to Table 14 Appendix for the exothermic equation and rate for $H_2O(g)$ to emphasize use of state symbols in thermochemical equations. For endothermic reaction, write the two forms of thermochemical equations for the decomposition of H_2O. Relate to HRW Transparency, CH-7163 (B) for Fig. 19.1B. **Demonstration or films:** Either demonstrate the reactions or use the films: YF-241, *Slow Reaction—Iron and Oxygen*; YF-242, *Fast Reaction—Mercury and Silver Nitrate*; YF-243, *Very Fast Reaction—Ammonium Dichromate.*

Sections 19.3–19.4. **Activities:** Use the previous activities for reaction of $H_2 + O_2$ for developing the definition of heat of formation and the use of ΔH_f. Further discussion of energies involved in bond breaking and bond making is useful. Base the stability concept on the negative sign and the magnitudes of heats of formation. Have students select the compounds formed by endothermic processes from Table 19-1. Supplement by additional data for nitrogen tri-iodide, hydrogen sulfide, and mercury fulminate. Refer students to Table 14 Appendix for identifying types of compounds as very stable, generally unstable, explosive. Have students discover the relationship between stability and electronegativity difference for pairs of compounds, i.e. NaCl-HCl, HCl-HI. **Transparencies:** HM 2-10540, "Heat of Reaction—Enthalpy Change"; C T 600-51, "Energy and Chemical Reactions"; C T 605-9, "Heats of Formation." **Filmstrip:** DG 6501, "Thermochemistry" and related transparency set DG TC-2.

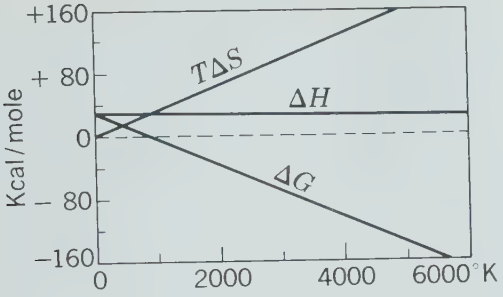
Section 19.5. **Demonstration:** (a) Burn a charcoal stick in a bottle of O_2, test for CO_2 with limewater. Use as lead-in to heat of combustion (formation of CO_2) of C (graphite). Write the equation using ΔH, define heat of combustion, and use ΔH_c. Ask students why the heat of formation of carbon dioxide is the same as the heat of combustion of carbon. (b) Burn CO and identify CO_2 product. Show how the heat of formation of CO is calculated, Sec. 19.5. This leads to the concept of conservation of the energy of a system during chemical activity (additivity of heats of reaction). Write the thermochemical equations for the two demonstrations (above). Have students recall that carbon is a reducing agent (reduction of CuO in Experiment 31, "Carbon") to suggest the problem of direct measurement of the heat of formation of CO. Have them suggest a method for its calculation from the data. Sequentially use the equation series in Sec. 19.5 and arrive at the principle to the right. Show how the heat of formation of CO_2 may be calculated from the heat of formation of CO and the heat of combustion of CO, using the equation series in Sec. 19.5. This example illustrates the principle of additivity of heats of reaction. **Transparency:** HRW CH-7164 (Fig. 19-2). To illustrate the additivity principle. (Hess's law: the heat of a given overall reaction is the same regardless of the intermediate steps involved). Use the series of equations for the calculation of ΔH_f for CO.



Section 19.7. **Review:** Concepts in Sec. 2.15, Reaction tendencies. Use the HRW transparency, CH-7106 (Fig. 2-9) "Change of Phase." **Demonstrations optional:** From Chapter 2 on energy-entropy relationships as they relate to physical changes and spontaneous processes: evaporation of volatile liquids, "disappearance" of Dry Ice, using the Hg-glass molecular vibration tube (Welch 1924). **Demonstrations:** The following lead to the lesson problem of spontaneous endothermic water gas reaction. (a) Tendency to lowest possible energy state: (1.) Zn + HCl. Note exothermic nature by feeling outside of test tube or reading inserted thermometer, (2.) Mix H_2O and coke, pass steam over coke, demonstrate water gas production: Alyea, pp. 9, 97, (p. 142 apparatus for Burning Magnesium in Steam may be adapted, substituting and preheating coke). Fowles pp. 113, 503, 571. (b) Tendency to highest possible entropy state:

(1) Simulate Fig. 19-4 using a colored gas (NO_2, Br_2) in one vessel and air in the other. Open connecting valve. Note result. Keep demonstration set up displayed for a considerable time. Note that gases do not reorder themselves. For free-energy change, define free energy change as net driving force. Write the equation, identifying each symbol in the equation. **Develop the principle:** A chemical change tends to proceed spontaneously in the direction of diminished free energy content (ΔG is negative). Write the thermochemical equation for the water gas reaction and have students attempt to explain the spontaneity of the reaction at elevated temperature. Perform the calculations (Sec. 19.7) determining the ΔG for the water gas reaction at 25° C. Repeat for

1000° K at which the free energy change becomes negative.



"JANAF Thermochemical Tables" The Dow Chemical Co. Midland, MI, 1961.

Transparency: HRW CH-7126, Molecular speed distribution in a gas at different temperature, may be useful. See the diagram from *Chemical Energy* above. It helps explain the spontaneity of the water gas reaction at the higher temperature by showing the variations of ΔG and $T\Delta S$ with temperature. **Reference:** See Toon and Ellis, *Foundations of Chemistry*, pp. 427–428 for summation of thermal effects on H and S, and spontaneity of reactions. **Films:** ROB, *Entropy*; ROB, *The Second Law*; ROB, *Patterns of Change*; ROB, *Molecules at Work*; ROB, *Times of Change*. **Filmstrip:** DG 679, "Energy and Entropy, The Driving Force of a Chemical Reaction." **Questions:** 6, 7, 15–18.

(2) Spontaneous endothermic reactions:

Matthews, G. W. J., *J. Chem. Educ.*, 43, 476 (1966). Measurable temperature decrease and color changes as 1 g of a transition metal hydrated chloride such as $CoCl_2 \cdot 6H_2O$ reacts with 3–5 ml of thienyl chloride. Increase in entropy as large volumes of gases are released. $MCl_2 \cdot 6H_2O + 6SOCl_2 \rightarrow MCl_2 + 12HCl(g) + 6SO_2(g)$.

(3) Free energy:

Alyea, H. N. and Dutton, F. B., *Tested Demonstrations in Chemistry*, pp. 221–223. Eberhardt, W. H., "Concerning Equilibrium, Free Energy Changes, Le Chatelier's Principle, I, II, III."

I. $Cu^{++} + 2I^- = CuI + \frac{1}{2}I_2$

II. $CoCl_4^{--} + 6H_2O \rightleftarrows Co(H_2O)_6^{++} + 4Cl^-$

III. Halite-Halate Equilibria

(A) KI and KIO_3 (C) KBr and $KBrO_3$

(B) KI and $KBrO_3$ (D) KBr and KIO_3

$5Br^- + BrO_3^- + 6H^+ \rightarrow 3Br_2 + 3H_2O$

(4) Enthalpy decrease versus entropy decrease: Mix concentrated solutions of $AgNO_3$ and either HCl or NaCl at room temperature. Inserted thermometer registers temperature increase. Precipitation of AgCl indicates entropy decrease. Difference between $\Delta H = -15.7$ kcal/mole and $\Delta G = -13.2$ kcal/mole is the "organization" energy required for the ordering of the crystalline AgCl.

(5) Probability and entropy. Recommended references for teachers and for interested students include: Young, J. A., *Chemistry: Reflections of Another Teacher*, Silver Burdett, 1965; Ford, K. W., "Probability and Entropy in Thermodynamics," *The Physics Teacher*, 5, 77, 1967.

Section 19.8. **Activity:** Use colored polystyrene spheres held together by magnets to simulate the pathways or possible pathways and the representative equations for the HBr and HI reactions. **Student references:** on flash photolysis, *Chemistry* 40 (1967) "Flash Photolysis," Sept.; "Bewhiskered Theory Disproved," (the HI reaction), Apr. *Scientific American*, 219 (1968) "The Chemical Effects of Light," Sept., (Flash Photolysis and the photolytic reaction for HCl). **Demonstration:** (a) Photochemical Reaction of Hydrogen and Chlorine, Alyea, p. 154. Plastic pharmaceutical vials or "baggies" may be substituted for the balloons mentioned in the reference. Clock reactions: Alyea, pp. 19, 85, 130, 147, 149, 179, 220. Clock reactions are interpreted in terms of the steps involved. (b) Chen, P. S., *Journal of Chemical Education*, 47 A784 (1970): A reaction between 100 ml of 1-M solution of $FeCl_3$ mixed with 100 ml of 1-M $Na_2S_2O_3$ results in three color changes which are explained in the reference. **Filmstrip:** DG 6500, "Chemical Kinetics," corresponding transparency set DG TC-1. A part of these deals with the HCl mechanism.

Sections 19.9–19.10. **Transparencies and activities:** Transparency, HRW CH-7126 corresponds to Fig. 19.6. Use models to simulate collision patterns in Fig. 19.7 (HRW transparency 7169), emphasizing the necessary violence (energy) and orientation in terms of activation energy, bond breaking, bond making, and electron cloud interpenetration. Toon and Ellis, *Foundation of Chemistry*, Fig. 12-7 and p. 385 shows how the number of possible collisions (collision frequency) depends on the number of molecules present. Use the term collision effectiveness with reference to sufficient energy and orientation.

Section 19.11. Use the conversion of kinetic energy into potential energy idea at beginning of Sec. 19.11. Place a vertical line representative of required activation energy on diagram Fig. 19-6 or the transparency CH-7126. This shows how increase in the number of collisions with increasing kinetic energy (temperature) increases the efficiency of collisions since a large portion of the molecules reach activation energy. Use Fig. 19-11 (HRW CH-7168) and Fig. 19-12 (HRW CH-7169) in conjunction with models. Then use HRW transparency CH-7170 for discussion of Fig. 19-13, for energies in the HI reaction. **Films:** CHEM 4121, *Introduction to Reaction Kinetics*; CHEM 4166, *Mechanism of an Organic Reaction*; MH *Energy and Reaction.* **Loop:** CHEM 168W1012 or 168W1031, "An Organic Mechanism. **Filmstrips:** HR C4-96455 *Reaction Kinetics*, DG 6500 *Chemical Kinetics.*

Section 19.12. **Laboratory Experiments:** Experiment 35, Rate of Chemical Reaction; Experiment 36,

"Course of a Chemical Reaction: Catalysis." Develop the lesson on the results of Experiment 35. Experiments from *Scientific Experiments in Chemistry* are: Experiment 39, "Rate of Hydrolysis of an Ester"; Experiment 41, "Effect of Concentration on Rate of Reaction"; Experiment 42, "Concentration of Hydrogen Peroxide and its Decomposition Rate"; Experiment 43, "Effect of Temperature on Rate of Reaction"; Experiment 44, "Effect of Weight of Catalyst on Reaction Rate."

Films:

MLA CHEM 4121 *Introduction to Reaction Kinetics* (16 mm, 13 min, c, s)
MLA CHEM 4166 *Mechanism of an Organic Reaction* (16 mm, 20 min, c, s)
MLA CHEM *Introduction to "Reaction Kinetics"* (16 mm, 16 min, b/w, s)
MLA CHEM 4127 *Catalysis* (16 mm, 17 min, c, s; 8 also available #4827)
MLA CHEM 4027 *Teacher Training Introduction to "Catalysis"* (16 mm, 13 min, b/w, s)
MLA CHEM *Introduction to "Mechanism of an Organic Reaction"* (16 mm, 9 min, b/w, s)
FAC *Chemical Change and Temperature* (16 mm, 14+ min, c, s)
MH 612011 *Energy and Reaction* (16 mm, 14+ min, b/w, c, s)
FAC 16-253 *Speed of Chemical Change* (16 mm, 15 min, c, s)
EBF 253 *Velocity of Chemical Reactions* (16 mm, 11 min, b/w, s)
ROB *Entropy* (16 mm, 30 min, b/w, s)
ROB *The Second Law* (16 mm, 30 min, b/w, s)
ROB *Patterns of Change* (16 mm, 30 min, b/w, s)
ROB *Molecules at Work* (16 mm, 30 min, b/w, s)
ROB *Times of Change* (16 mm, 30 min, b/w, s)
AIM Yale Films (16 mm, s)
YF-241 *Slow Reaction—Iron and Oxygen* (2 min, c)
YF-242 *Fast Reaction—Mercury and Silver Nitrate* (5 min, c)
YF-243 *Very Fast Reaction—Ammonium Dichromate* (1 min, c)
YF-223 *The Bomb Calorimeter* (9 min, c)
YF-214 *Catalytic Decomposition of Hydrogen Peroxide* (2 min, c, s)
YF-215 *Catalytic Decomposition of Potassium Chlorate* (2 min, c, s)
SUTH *Rate of Combustion* (8, 3 min, c, s)

Loops:

MLA CHEM 168 W 1011(T) or 168 W 1031(K) *An Organic Mechanism* (s-8, 4 min, c)
H-E or E E-84-0231/1 *Reaction Rates* (4 min, c)
W *Reaction Rates* (s-8 T or s-8 K, c, 4 min)
HR 04-96455 *Reaction Kinetics* (s-8, 4 min, c)

Filmstrips:

DG 6500 "Chemical Kinetics" (40f, c)
DG 6501 "Thermochemistry" (40f, c)
DG 679 "Energy and Entropy: The Driving Force of a Chemical Reaction" (40f, c)
DG 689 "Bond Energies" (40f, c)

Transparencies:

HRW CH-7163 Change in heat content during chemical action
HRW CH-7164 Heat of formation diagram for CO and CO_2
HRW CH-7165 A possible mechanism for the water-gas reaction
HRW CH-7166 Pathway of an exothermic reaction
HRW CH-7167 Pathway of an endothermic reaction
HRW CH-7168 Activation energies for the forward and reverse reactions
HRW CH-7169 Possible collision patterns for HI molecules
HRW CH-7170 Possible activated complex configuration which could form either 2HI or H_2 plus 2I
HRW CH-7171 An energy profile of hydrogen iodide
HRW CH-7172 Possible difference in potential energy change along alternate reaction pathways
HRW CH-7173 Reaction rates for the hypothetical reaction system
K & E 30-6120-37 Rates of Reaction—Effects of Catalysts (1 + 9)
K & E 30-6120-38 Factors Affecting Reaction Rates (2 + 2)
HM 2-10540 Heat of Reactions—Enthalpy Change
HM 2-10541 Catalysis and Activation Energy
HM 2-10542 Molecular Distribution of Kinetic Energy
C T-600-51 Energy and Chemical Reactions
C T-600-52 Energy of Activation—Exo- and Endothermic and Catalyzed Chemical Reactions
C T-605-9 Heats of Formation
McW UNIT IX Reaction Rate Kinetics
Energy Distribution Curve
Potential Energy Diagram for Endothermic Reaction
Potential Energy Diagram for Exothermic Reaction
Collision Geometry
DG Set TC-1 Chemical Kinetics (7 transparencies with 14 overlays)
TC-2 Thermochemistry (6 transparencies, 12 overlays)

TEACHER NOTES CHAPTER 20

Chapter Note: Activities and Demonstrations: Students should perform as many as possible of the suggested experiments. Experiments not scheduled for the class may be used as student projects. The results of these are reported by the students to the class. Students may also perform some of the suggested demonstrations for the class.

Section 20.1. Demonstration: Have students note that the hot black residue of the heated mercury (II) oxide (HgO) changes back to the orange color of HgO. Mention the work of Priestley and Lavoisier and the reversibility of the reaction. Discuss how, in a closed system, equilibrium can be reached. You may wish to use the data from Sec. 19.7 ($\triangle S = 0.032 \frac{\text{kcal}}{\text{mole} \cdot ^\circ\text{K}}$; $\triangle H = +31.39 \frac{\text{kcal}}{\text{mole}}$) to calculate

the temperature at which the water = gas system reaches equilibrium. By substituting in the equilibrium equation $T\triangle S = \triangle H$, which represents a balance of the driving forces (entropy change = energy change), the equilibrium temperature is calculated to be approximately 980° K. Demonstration: one or both aspects of reversible reaction in Fig. 20-1. To demonstrate hydrogen production, place a piece of asbestos sheet saturated with H_2O into the bottom of a test tube. A wad of steel wool, positioned in the middle of the horizontal test tube, is preheated. The delivery tube from the reaction test tube is connected to a test tube arranged for water displacement. Heat the bottom of the test tube to vaporize the water and simultaneously heat the steel wool. Identify the collected H_2 by the bark test. To demonstrate reduction use Fe_2O_3 instead of Fe_3O_4 since the former undergoes a perceptible color change and the iron produced is magnetic while the magnetic iron oxide (Fe_2O_3) is not. Reversibility demonstration: By heating, convert hydrated copper (II) sulfate $CuSO_4 \cdot 5H_2O$ crystal to white anhydrous powder, $CuSO_4$. Add the powder to water. Add a few drops of water to the powder.

Section 20.2. Demonstration: Repeat equilibrium vapor pressure demonstrations from Chapter 11 to review physical equilibrium and its dynamic nature. Discussion: of equilibrium vapor pressure (and saturated solutions) reveals that for a given closed system, the particular equilibrium state depends upon temperature and, for systems involving gases, also depends on the pressure. Demonstration: Demonstrate saturated solutions of NaCl and sugar. To indicate that an equilibrium state may be recognized by observation of a macroscopic property, drop a crystal of NaCl (rock salt) into a filtered portion of the saturated solution which was prepared before the class. Write ionic equations representative of ionic equilibria for NaCl(s) and $HC_2H_3O_2$. Demonstration: Repeat conductivity test of glacial acetic acid. Test after several successive dilutions. This shows concentration to be a factor in equilibrium states. To show reversibility of this reaction, add H_2SO_4 to solid $NaC_2H_3O_2$. Have students detect the odor of acetic acid or test for its release by moistened blue litmus paper in the upper part of the test tube. Discussions: Discuss the ionization of diluted acetic acid as an example of a reaction reaching equilibrium far to the left since acetic acid is a weak acid (test with pH paper if necessary). Discuss the dissociation of NaCl or the ionization of HCl as reactions that go far to the right. Discuss the chemist's problem of producing desirable products in terms of the ammonia synthesis. Have students restate the five factors that determine the rate of chemical reaction.

Section 20.3. Discussion: Develop Fig. 20-2 on the board, starting with zero concentrations of C and D. Discuss the changes in rates as concentrations change. Define the equilibrium constant. Write the representative equation, and the equilibrium constant expression for $A + B \rightleftarrows C + D$ and for its more general form. Use the terms chemical equilibrium law and mass action expression. Have students apply these terms to the equation $3A + B \rightleftarrows 2C + 3D$ or variations of this equation. Derive the expression for the HI reaction. Discuss constancy of color of this system as a macroscopic method of detecting equilibrium. Problem: Solve the Sample Problem for Sec. 20.3. Divide the class into four sections and have each section calculate the K from separate experimental data listed in Table 20-1. Follow through the sequences of equations and related expressions for the rate laws which lead to K. Use for the one-step and two-step reactions. This should lead to the conclusion that the expression for K is the same whether calculated from the balanced chemical equation or from the rate law treatment. Questions: 1-5, 11, 12, 16, 21.

Section 20.4. Discussion: Raise the problem of how chemists may displace an equilibrium in order to increase desired production. Review and restate Le Chatelier's principle (Sec. 11.5) and the related demonstrations relating to physical changes such as regeletion (demonstrated by squeezing two ice cubes together) and using film YF-225, *Le Chatelier's Principle*. Also use YF-250, *Vapor Pressure*. These films reveal that pressure and temperature changes affect equilibrium. Laboratory experiments: Experiment 37, "Equilibrium;" Experiment 38, "Solubility-Product Constants;" Experiment 39, "Hydrolysis;" Experiment 40, "Equilibrium and Complex Ions." From *Scientific Experiments in Chemistry*: Experiment 45, "Solubility-Product Constant of Cadmium Hydroxide," is an alternate to Experiment 41. Reference: A recommended reference is Morris, D. L., "Stress, Collisions and Constants," *Chemistry*, 44 (1971) Apr., which is Part I in the readily available reprint 100. This article complements the text by treating various factors from the viewpoint of Le Chatelier's principle, collision theory, and the equilibrium constant. Demonstrations: Factors that disturb equilibrium—(a) Change in concentration: (Some of the demonstrations here also fit Sec. 20.6 Common Ion Effect and Sec. 20.12 Solubility Product. Some of the suggested demonstrations for those sections may be used here also.) 1. Review or demonstrate Experiment 37, Equilibrium; use procedures 1-4 and "Further Experimentation" which describes the precipitation of NaCl from its saturated solution by (a) HCl and (b) (NaOH.) 2. Introduction to Experiment 38, "Solubility-Product Constants," discusses the dichromate-chromate equilibrium and the effects of H_3O^+ and OH^- addition. 3. See Experiment 40, "Equilibrium and Complex Ions" for concentration effects on complex ions of Cu (II) and Co(II). Demonstrate addition of concentrated HCl to solution of $CoCl_2$. 4. References: *Journal of Chemical Education*, 47, A438 (1970) Dems. 787, 788, 790. Summarize results of selected demonstrations by using the text development Sec. 20.4.1. (b) *Change in pressure:* 1. Instructor's Experiment in Experiment 39 describes effect of decreased pressure by vacuum on

pH of a $NaHCO_3$ solution. 2. See Campbell, J. A., *Journal of Chemical Education*, 47, 276 (1970), "sealed Tube Experiments," No. 35 describes rapid compression of NO_2 in a plastic syringe. The demonstration should probably follow the temperature effect on the $2NO_2 \rightleftharpoons N_2O_4$ system; the initial darkening due to heat of compression eventually leads to the expected fading of color. 3. Homogeneous system: Discuss the Haber process with the aid of HRW transparency CH-7174 (Fig. 20-3). Consider the effects of continuous removal of NH_3 by condensation. Discuss pressure effect on $CO(g) + H_2O(g) \rightleftharpoons CO_2(g) + H_2(g)$. 4. Heterogeneous reactions: Develop K for the decomposition of $CaCO_3$ and discuss treatment of pure solids (and liquids) in the expression for K and effect of pressure changes on K. (c) *Temperature changes:* Discuss the effects of temperature increase on the rates of both reactions, the favoring of the endothermic reaction according to the Le Chatelier principle, and the effect on the equilibrium constant. Relate to the commercial production of ammonia. Use the data in Toon and Ellis, *Foundations of Chemistry*, Table 13-1, p. 411 and read pp. 410–412. Demonstrate effects of temperature changes on the $2NO_2(g) \rightleftharpoons N_2O_4(g)$ system, $\Delta H = 14.1$ kcal (Welch 4226 Temperature Equilibrium Tubes). Sealed test tubes of NO_2 may be homemade (See *Chemistry*, 45, 1972, Jan., "Making Tubes for NO_2–N_2O_4 is practically complete. See Experiment 40, Instructor Experiment for temperature effects on saturated solution of $CoCl_2$; The deep red solution turns blue upon heating. Placing the bottom of the heated tube into an ice-salt mixture causes it to turn back to red. Use the data in Table 20-2, for the Haber process to show that by changing temperature, the value of the equilibrium constant changes. Refer also to HI data. Consider the composition mixture when the K is (a) large, (b) very small. Demonstrate the $Cu + Ag^+$ replacement for large K. Refer to Table 20-4 Solubility-Product Constants for very small value for K and precipitate AgCl to demonstrate slight solubility of AgCl in a saturated solution. Procedure 2, Experiment 36 uses more soluble silver acetate. Also see TOPS Dem 787, Equilibrium AgOAc, *Journal of Chemical Education*, 47 A 437 (1970). (d) *Catalysts:* To show that a catalyst accelerates both the forward and reverse reaction (and does not affect the value of K) demonstrate decomposition of dry NH_3. Pass NH_3 over heated steel wool. Collect the H_2 and N_2 mixture by displacement of water. Obtain a postiive "bark" test for H_2. Consult the listed Alyea references for other demonstrations, e.g., "Introduction to Le Chatelier's Principle," p. 194. This shows that the solution of dry HCl in H_2O is exothermic and hence heating concentrated hydrochloric acid reverses the solubility.

Section 20.5. Demonstration: Demonstrate in sequential fashion the three possible situations which are representative of reactions that run to completion. Many teachers prefer to start with neutralization, the formation of slightly ionized products, then proceed to volatility and finally to precipitation. This sequence aids students to determine which reactions run to completion. The reaction is neutralization if the reactants are an acid and a base. Secondly, the reactants are examined to determine if a gas is a possible product. Finally the students know a precipitate forms by the use of solubility rules and/or tables to check on possible products. The first step is writing the formula equation. This equation is then converted to the ionic form and the net ionic equation is finally written after checking the equation for the various potentialities. Demonstrations: As the three possible situations are being established, demonstrate the reactions suggested in the text. Propose situations such as mixing solutions of $CuSO_4$ and NaCl or NaCl and KNO_3. Perform these as demonstrations and have students reach conclusions. Demonstrate the reaction between $Ba(OH)_2$ and H_2SO_4. Directions are in Alyea, p. 163. This reaction is the subject of the film, YF-236, *Removal of Ions by Metathesis*. Both indicator and conductivity apparatus are used and the conductivity is reduced to zero at the neutralization end point since ions are removed by precipitation and water formation. Another variation of this demonstration uses a small beaker containing 50 ml of 0.1-*N*.Na_2SO_4. Two platinum electrodes reach into this solution. A thin layer of paraffin is poured over the solution and permitted to harden. A conductivity reading may be made by meter or bulb. Pour 50 ml of 0.1-*N* $BaCl_2$ solution on top of the paraffin. Another conductivity observation is made. The paraffin is broken, the solutions mixed and a reading taken. Here, as in the neutralization of NaOH with HCl, one set of ions is still left after the reaction has run to completion. Use solutions of $MgSO_4$ and $Ba(OH)_2$ with the set-up. The precipitation of two compounds $Mg(OH)_2$ and $BaSO_4$ takes place. Questions 8, 9, 17.

Section 20.6. Demonstration: Add solid $NaC_2H_3O_2$ to a dilute solution of $HC_2H_3O_2$ containing either methyl orange (pink-red changes to orange) or methyl red (violet changes to red). Use pH paper or meter to measure pH effect. Explain effect as an application of Le Chatelier's principle. Add concentrated HCl to a filtered, saturated solution of NaCl and observe precipitation of NaCl. Try addition of ten pellets of NaOH to the saturated NaCl solution. Refer to Alyea: p. 19, NH_4Cl to NH_3(aq) with phenolphthalein; pp. 86, 87; pp. 198, 200, to compare foam produced by acetic acid and $CaCO_3$ versus that produced by sodium acetate + acetic acid on $CaCO_3$); p. 210, to compare the rate of balloon filling from $Mg + HC_2H_3O_2$ versus from $Mg + HC_2H_3O_2 + NaC_2H_3O_2$.

Section 20.7. Activities: Recall or redemonstrate conductivity tests for acetic acid (glacial and diluted) and for water using neon or argon bulb. Use the text development for the derivation of K and then K_a for acetic acid. Have students calculate the K_a using the data in Table 20-3. Solve the related text problem on

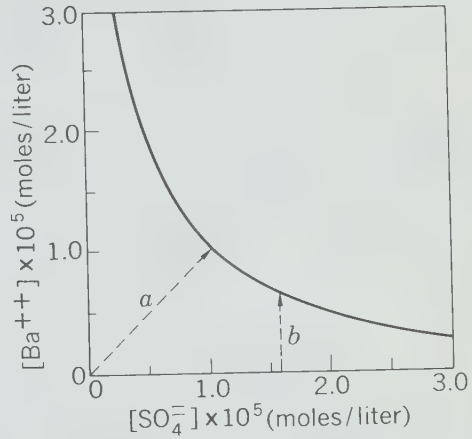
pH data. Use the equation and K_a expression for the general form of a weak acid ionization. Discuss the effect of addition of common ion on the hydronium ion concentration and the effect of temperature increase on K_a according to the Le Chatelier principle. Demonstrations: Demonstrate buffering action by adding a small quantity of NaOH solution to a solution of acetic acid and sodium acetate using methyl red or methyl orange indicator.

Section 20.8. Discussion: Derive the expression for K and K_w for water as in Sec. 20.8. Define ion-product constant for water. Review previous treatment in Sec. 15.4 and of pH in Sec. 15.5 and Sec. 15.6. Recall the changes of K_w with temperature change, e.g., values at 0° C and 60° C.

Section 20.9. Activity: Use student results from Experiment 39, "Hydrolysis." Exhibit pharmaceutical box of $NaHCO_3$ and commercial box of $Na_2CO_3 \cdot 10H_2O$. Ask students why baking soda may be safely ingested as an antacid while the Na_2CO_3 can not. Test separate solutions of $NaHCO_3$ and Na_2CO_3 made from solids in front of class by simultaneously inserting and quickly withdrawing red litmus paper (the $NaHCO_3$ gives a much lighter blue than the Na_2CO_3). Follow up with Hydrion pH paper tests. Ask why washing soda is widely used as a cleaning agent and refer to pH table of Common Substances (Chap. 15). Discussion: Define hydrolysis. Discuss, illustrate and, if necessary, demonstrate the two general types. Use the general equation for basic anion hydrolysis and the expression for K_h as in Sec. 20.10.

Section 20.12. Discussion: Open the discussion with a question pertaining to saturated solutions and their qualitative descriptions. Refer students to Table 13, Appendix, Solubility Chart, for examples of substances whose saturated solution would be (1) concentrated or (2) dilute. Demonstration: Add solid $CaSO_4$ to water. Test the filtrate with the flame test for calcium and the test for sulfate ion. Review terms associated with solubility: soluble, slightly soluble, partly soluble, sparingly soluble, and insoluble. Select examples of soluble, insoluble, and slightly soluble from Appendix Table 12. Precipitate $CaSO_4$ by mixing solutions of $CaCl_2$ and Na_2SO_4. Use results of Experiment 38, "Solubility Product Constants," and from *Scientific Experiments in Chemistry:* Experiment 21, "Progressive Precipitation," and Experiment 45, "Solubility Product Constant of Cadmium Hydroxide." Use the text examples (AgCl and CaF_2) for deriving the K and K_{sp} for the systems and define solubility product constant. Go through the calculations of K_{sp} for each of these salts. Have students refer to Table 13, Appendix, for AgCl used in the calculation of its K_{sp}. Demonstration: The article by Morris, D. L., *Chemistry*, 44, Apr. (1971); "Stress, Collisions, and Constants, Part I, solutions," suggests a method for using NaCl in dealing with K_{sp} and the related calculations.

Section 20.13. Use the text example and emphasize the resultant ion concentrations. As an application have students solve the reversed CaF_2 problem in Sec. 20.12. In Sec. 20.14, Precipitation calculations, reemphasize the concept that K_{sp} is an equilibrium constant and that therefore varying ion concentrations may result in exceeding the K_{sp}. Solve Problem in Sec. 20.13. Use the portion of the DG 6502 filmstrip *Le Chatelier's Principle* that deals with the equilibrium states of $BaSO_4$. Also use the portion of the related transparency set DG TC-3 that covers the BaSO K_{sp}



Bruce H. Mahan, UNIVERSITY CHEMISTRY, SECOND EDITION, 1969, Addison-Wesley, Reading, Mass.

The diagram represents the equilibrium states of the $BaSO_4$–H_2O system. The equilibrium states of the system lie on a rectangular hyperbola. The broken line *a* represents the equal increase in concentration of both Ba^{++} and SO_4 as $BaSO_4$ dissolves. The broken line *b* represents the increase in concentration of Ba^{++}, if small amounts of $BaCl_2$ are added to a solution of H_2SO_4. At the point where *b* intersects the hyperbola the first $BaSO_4$ precipitates. As more Ba^{++} is added, more $BaSO_4$ precipitates and the $[SO_4^{=}]$ decreases. As this process continues the point representing the system follows the curved line toward decreasing $[SO_4^{-}]$. Thus, any point on the curved line represents one of the many possible equilibrium states of the system. Use questions 9, 10, and problems 8, 9, 10. For additional problems, see Castka, *Chemistry Problems*, Chap. 9. Another source is Toon and Ellis, *Foundations of Chemistry*, Chaps. 13 and 16. Use the references for programmed instruction at the end of the TE chapter for students who can profit from such activity.

Films:

MLA CHEM 4124 *Equilibrium* (16 mm, 24 min, c, s)

MLA CHEM 4024 *Teacher Training Introduction to "Equilibrium."* (16 mm, 7 min, b/w, s)

ROB *Equilibrium—The Limit of Disorder* (16 mm, 30 min, b/w, s)

Note: See other ROB films in Chap. 19, TE.

INDU FS-427 (b/w), FSC-427 (c) *Ionization and Ionic Equilibrium* (16 mm, 15 min, s)
AIM Yale films
YF-216 *Decomposition of Mercuric Oxide* (16 mm, 3 min, c, s)
YF-237 *Solubility Product* (16 mm, 7 min, c, s)
YF-250 *Vapor Pressure* (16 mm, 7 min, c, s)
YF-225 *Le Chatelier's Principle* (16 mm, 3 min, c, silent)
YF-233 *Triple Point Determination* (16 mm, 5 min, c, s)
YF-236 *Ion Removal by Metathesis* (16 mm, 4 min, c, s)

Loops:
H-E or E E-84-0223/1 *Equilibrium: Le Chatelier's Principle*, (s-8, 3 min, 30 sec, c)
H-E or E E-84-0215/1 *Equilibrium: Solutions and Precipitates*, (s-8, 3 min, c)
HR *Fundamentals of Chemistry, Buffer Solutions* (s-8, c)

Transparencies:
HRW CH-7174 Pressure increases the yield of Ammonia
HRW CH-7175 The Heterogeneous Equilibrium in a saturated solution
DG TC-3 Le Chatelier's Principle (4 + 8)
DG TC-4 Reversible Reactions and Reactions to Completion (5 + 10).
HR 2-10543 The Law of Chemical Equilibrium.
HR 2-10544 Factors Affecting Chemical Equilibrium
K & E 30-6120-34 Reactions that go to completion (1 + 6)
K & E 30-6120-39 Equilibrium Vapor Pressure (2 + 7)
K & E 30-6120-40 Equilibrium in a Saturated Solution (1 + 2)
K & E 30-6120-41 Concentration of Solution (1 + 4)
K & E 30-6120-43, 44, Chemical Equilibrium Affecting Factors Part 1 (1 + 3), Part 2 (5)
K & E 30-6120-45 Common Ion Effect (1 + 5)
K & E 30-6120-71 Hydrolysis of Dissolved Salts (1 + 6)
C T-600-20 Le Chatelier's Principle Applications
C T-600-41 Variation of Vapor Pressure with Temperature
C T-599-55 Equilibrium in Reversible Chemical Reactions
C T600--56 Equilibrium Constant
C T-600-15, 16 Chemical Equilibrium

Filmstrips:
DG 658 Equilibrium Constant (40 f, c)
DG 6502 Le Chatelier's Principle (40 f, c)
DG 6503 Reversible Reactions and Reactions to Completion (40 f, c)
DG 692 Indicators (40 f, c)

TEACHER NOTES CHAPTER 21

Chapter Notes: Activities and Demonstrations: Students should perform as many of the experiments related to oxidation-reduction as possible. These include: Experiment 41, "Oxidation-Reduction Reactions;" Experiment 42, "Oxidation-Reduction Titration;" Experiment 43(A), "Electrochemical Cells, Half-cell Reactions;" and Experiment 46, "Electrolytic Cells." The rules for determining oxidation numbers are reviewed and used extensively in balancing redox equations. The demonstrations suggested and the numerous audio-visual aids available provide for a choice of those activities that are best suited to the abilities of your students and the requirements of your course. A film which may be used as an introduction or a summary is IBF 393 *The Development of Electrochemistry*.

Section 21.1. Demonstrations: If necessary, repeat previous demonstrations relating to oxidation and reduction primarily in Chap. 6. Sequentially these are: (a) *Oxidation-reduction involving oxygen*: oxidize a strip of copper and reduce copper(II) oxide (strip or pellets) with H_2, Experiment 9, *Scientific Experiments in Chemistry*. Copper(II) oxide is reduced by carbon, Procedure 7, Experiment 31, "Carbon." (b) *Oxidation-reduction involving electron shift:* burn Na in Cl_2 (Alyea pp. 146, 201, 212). Burn Mg in air and in Cl_2 (*caution*). Use oxidation electronic equations in the form $Na \rightarrow Na^+ + e^-$ for half-reactions. Burn steel wool in air and Cl_2. Sprinkle antimony into Cl_2. (c) *Oxidation-reduction for covalent species:* burn H_2 in Cl_2, Procedure 4, Experiment 61, "Chlorine." Define oxidation and reduction. Explain their simultaneous occurrence. Apply oxidation rules to text examples and questions 6, 7. By using text examples and equation series, emphasize that oxidation results in algebraic decrease in oxidation number of a species while the reverse is true in reduction. Use questions 1-3 and 4 to find oxidation-reduction reactions from among a list of equations.

Section 21.3. Discussion and questions: State and illustrate the rules sequentially using the two text examples and the Sample Problem (Sec. 21.3). Use questions 11-15. Transparency: C T 600-26 "Oxidation-reduction-Balancing of Equations by the Oxidation Number Method." Reference articles: *Chemistry:* "Balancing Oxidation-Reduction Equations," 37, July (1964); "How Do Redox Reactions Take Place?" reprint 20, 38, Mar. (1965); "Redox Reactions—Why Complicate the Subject?" 41, July-Aug. (1968). Activity: You may wish to use examples of balancing equations in ionic form, e.g.,

$$H^+ + Cl^- + MnO_4 \rightarrow H_2O + Mn^{++} + Cl_2$$

$$Cu + H^+ + NO_3^- \rightarrow Cu^{++} + H_2O + NO$$

Check the correctness of the result by employing the principle of conservation of ionic charge. Compare the resultant ionic equation with the balanced formula equation. Various methods for balancing redox equations are illustrated in the filmstrip DG 642, "Oxidation-Reduction." **Laboratory Experiments:** 41, "Oxidation-Reduction Reactions"; Experiment 42, "Oxidation-Reduction Titration."

Section 21.4. **Discussion and Transparency:** Define oxidizing and reducing agents. Develop the relationships in Table 21-2 on the chalkboard. Discuss Group I elements as reducing agents and Group VII elements as oxidizing agent. Use HRW transparency CH-7125 (Fig. 6-10). Develop the marginal notes in Sec. 21.4. **Demonstrations:** Use the previous replacement of Cu^{++} by Zn (or equivalent) and of Br^- by Cl_2 to identify the respective half-reactions as oxidation or reduction. Use the loop HE 84-0140/1 *Oxidation-Reduction: the Halogens*, the CHEM 4112 film *Chemical Families*, and the CORF film *The Sodium Family*. Relate the demonstration reaction to Table 21-3. Some teachers use the generalization that a reducing agent will react spontaneously with any oxidizing agent below it to yield the weaker reducing agent and the weaker oxidizing agent (i.e., stronger + stronger → weaker + weaker). **Demonstrations:** (a) HCl + $KMnO_4$, HCl + $K_2Cr_2O_7$; identify oxidizing and reducing agents and write equations. (b) Hydrogen peroxide + catalyst (manganese dioxide, activated charcoal). Define autooxidation and write the equation (Sec. 21.4). Relate Table 21-3 to results of laboratory experiment 41.

Section 21.5. **Activities and demonstrations:** Define chemical equivalent in terms of the Avogadro number of electrons (a mole of electrons). **Demonstrations:** For iron to iron (II) ion use Fe + HCl. For iron to iron (III) ion burn steel wool in Cl_2. Students make the related calculations (Sec. 21.5). For iron (III) ion + tin (II) ion refer to results in Experiment 41, Procedure 4 or redemonstrate. Write half reaction equations, net ionic equation (redox) and make calculations as in Sec. 21.5. Discuss results of Experiment 42, "Oxidation-Reduction Titration" and calculations of equivalents.

Section 21.6. **Introductory demonstrations:** (a) "Spit and Eleven Cents." A silver dime (if available) and a penny are separated by a piece of blotting paper which is moistened with salt solution (some people use the title electrolyte). The leads from a lecture table galvanometer are touched to the coins. A piece of sterling silver may be substituted for the dime or try the existent devalued dime. (b) Use two dry cells in series to light a flashlight or use two beaker Zn-Cu-H_2SO_4 cells in which chemical action is evident (c) Dry cell equivalent: zinc strip, carbon rod or plates, blotting paper moistened with ammonium chloride or demonstration kit Welch 2242. Exhibit a dry cell cut in half. Precede the Sec. 21.6-1 dry cell treatment by demonstrating: 1. testing the temperature of the Zn + $CuSO_4$ or Fe (Steel wool is good) + $CuSO_4$ test tube reaction. 2. demonstrating electron transfer in external circuit of the equivalent cell (Zn-Cu-$CuSO_4$) by meter or deflection of a compass placed beneath or over the wire.

Section 21.6. **Demonstrations:** Electrolytic Cells. 1. To demonstrate that electrolysis is a non-spontaneous or driven reaction demonstrate that copper reacts spontaneously with silver (I) ion in an electrolytic cell; demonstrate that copper may be plated out on a silver electrode by supplying electric energy. The silver cathode is where the copper (II) ions form a solution of copper (II) nitrate takes place. Demonstrate that silver does not spontaneously replace copper in a solution containing copper (II) ions. 2. Use Fig. 21-4 or show CHEM 4109, *Electric Interactions in Chemistry*. To illustrate the equivalent of ion migration in an electrolytic cell use electrolysis of a dilute (blue) solution of copper (II) bromide. As copper plates out on the cathode, there is a lightening of the blue color. At the same time bromine is liberated at the anode and the blue color turns a dark green. Extract a sample of this greenish solution. Shake it with carbon tetrachloride or another solvent (liquid Freon). The characteristic orange-red color of bromine is produced. **Teacher Note:** We may define the cathode as the electron-rich electrode and the anode as the electron-poor electrode. Reduction occurs at the cathode and oxidation occurs at the anode in electrolytic (driven) cells. In electrochemical (spontaneous) cells the opposite reactions occur at the respective electrodes. This convention is reasonable since the direction of the current is reversed in the latter cells. In the storage or other rechargeable cells the same electrode always has the same name whether the cell is on charge or discharge cycle. With redox definitions, the names cathode and anode are usually dropped in the case of rechargeable cells because the names must change going from charge to discharge, and vice versa.

Section 21.7. **Demonstration and film:** Redemonstrate electrolysis of water. Use HRW transparency CH-7177. Explain in terms of equations and type of reaction at each electrode, the net cell equation, and changes in oxidation states. For details of electrolysis of water, using Na_2SO_4 see Further Experimentation, Experiment 43(B), "Electrolytic Cells." Use the film YF-230, Acid Base Reaction in Electrolysis of Water.

Section 21.8. **Demonstrate:** Electrolysis of NaCl solution using a U-tube with carbon electrodes. Provide a vent tube for the collection of H_2 at cathode and test for H_2 by "bark" test. Insert blue and red litmus papers in the solution at both electrodes. Hold them in place by the stoppers. Note the litmus changes at each electrode and have students explain. Have a student perform the waft test for Cl_2 by removing stopper

at anode. Write the electrode equations. Identify the processes. Write the net equation.

Section 21.9. Demonstrations: Refer back to any previous demonstrations of electroplating, i.e. copper on carbon electrode in copper (II) bromide. Demonstration series for Cu plating: (a) pass a low current (6v) through $CuSO_4$ using carbon electrodes until copper has plated. (Compare the adhesion of copper on the carbon rods to the deposit of copper in the reaction of zinc + copper sulfate.) Reverse the electrode polarity to remove the deposited copper. (b) "Electroplate copper onto a cleaned iron file by connecting the file with a copper electrode wire. Use no battery. Have students explain. (c) Use copper anode and carbon cathode. Reverse polarity after copper deposits out. (d) Use two copper electrodes and write electrode equations. Review discussion for naming electrodes (Sec. 20.9). Demonstrate silver plating. Contrast nature of deposit with that for test tube reaction ($Cu + AgNO_3$). Electrolyze a solution of HCl(2:1) using carbon electrodes and rather loose stoppers. Stop the electrolysis and immediately connect electrodes to a galvanometer. This procedure demonstrates the H_2–Cl_2 electrochemical cell. See Alyea, pp. 193, 213, 158. As summary use HRW transparency CH-7178 (Fig. 21-7).

Section 21.10. Demonstrations: Start with two lead plates in H_2SO_4 solution. Produce the storage cell by passing d.c. current through the set up until the production of brown deposit of PbO_2 is evident at the anode. Use the "charged" cell to ring a bell. Test for voltage of this cell with voltmeter. Use Fig. 21-9 as the equations are developed. Display a storage cell and demonstrate the use of a battery testing hydrometer.

Section 21.11. Demonstrations: Review Sec. 21.6 for processes occurring at each electrode in an electrochemical cell. See Experiment 43(A), "Electrochemical Cells, Half-cell Reactions." Set up the Zn-Cu cell (Fig. 21-10) and use the HRW transparency CH-7179. Demonstrate voltage measurement for this cell and establish half-cell equations. Define electrode potential. Exhibit a hydrogen electrode (Fig. 21-11); see Alyea, p. 141 "Electrode Potentials-Hydrogen Electrode" for construction.

Films:

MLA CHEM 4133 *Electrochemical Cells* (16 mm, 22 min, c, s; #4833: 8 mm)
MLA CHEM 4109 *Electric Interactions in Chemistry* (16 mm, 21 min, c, s)
MLA CHEM 4033 *Teacher Training Introduction to "Electrochemical Cells"* (16 mm, 8 min, b/w, s)
PSSC 0430 *EMF* (16 mm, 20 min, b/w, s)
PSSC 0431 *Electrical Potential Energy and Potential Difference* (Part I #0431, Part II #0432) (16 mm, 54 min total, b/w, s)
IFB 393 *The Development of Electrochemistry* (16 mm, 19 min, c, s)
SUTH *Oxidation-Reduction* (16 mm, 16 min, c, s)
MH 612009(b/w) 612021(c) *Faraday's Law* (16 mm, 15+ min, s)
MH 612001(b/w) 612013(c) *The Electromotive Force Series* (16 mm, 12 min, s)
EBF *The Primary Cell* (16 mm, 11 min, b/w, s)
AIM YF-230 *Acid-Base Reaction in Electrolysis of Water* (16 mm, 2 min, c)

Loops: H-E or E

E 84-0140/1 *Oxidation and Reduction: The Halogens* (s-8, 4 min, c)
E 84-0157/1 *Oxidation and Reduction: Electrolytic Cells* (s-8, 3.5 min, c)
E 84-0199/1 *Galvanic Cells: Electrolytes and Electrodes* (s-8, 3.5 min, c)
E 84-0207/1 *Galvanic Cells: Half-cell Reactions* (s-8, 4 min, c)
E 84-0181/1 *Hydrogen-Oxygen Battery and Fuel Cell* (s-8, 4 min, c)
H-E E 80-3940/1 *Production of Sodium: Electrolysis* (s-8, 4 min, c)

MLA CHEM STUDY LOOPS

168 W 1004(T) or 168 W 1024(K) *A Copper-Silver Electrochemical Cell* (s-8, 4 min, c)
168 W 1005(T) or 168 W 1025(K) *An Electrochemical Cell (animated mechanism)* (s-8, 4 min, c)
168 W 1013(T) or 168 W 1033(K) *Silver-Hydrogen Electrochemical Cell* (s-8, 4 min, c)
ICF 10100(8) or 10105(s-8) *Lead-Acid Storage Battery* (3 min, c)
ICF 10080(8) or 10085(s-8) *Voltaic Cells* (4 min, c).

Filmstrips:

DG 673 Fuel Cells (40 f, c)
DG 6504 Electrochemical and Electrolytic Cells (40 f, c)
DG 642 Oxidation-Reduction (40 f, c) Half-cell reactions, calculation of EMF, Balancing equations.

Transparencies:

HRW CH-7176 Diagram of a dry cell
HRW CH-7177 Electrolysis of water
HRW CH-7178 Comparison of electrochemical and electrolytic cells
CH-7179 A Zn-Cu voltaic cell
CH-7180 Measuring the oxidation potential of zinc half cell
R 21736 Electrolysis
R 21739 Galvanic Cells
K & E 6120 - series
K & E 6120 27 Oxidation-reduction (Change of Size) (1 + 4)
K & E 6120 60 Oxidation Numbers (1 + 2)
K & E 6120 62 An Experimental Voltaic Cell (1 + 8)
K & E 64 Electroplating (1 + 2)
C T-600 - series
C T-600-1 Activity Series
C T-600-24 Oxidation states (a)
C T-600-25 Oxidation States of Common Elements (b)
C T-600-26 Oxidation-Reduction-Balancing of Equations by the Oxidation Number Method (c)
C T-600-64 Electrolysis of Cupric Chloride Solution
C T-600-65 Oxidation Potential-Graphical Represen-

tations and Mechanical Addition of Half-Cell Voltages to Obtain Cell Voltage
T-605-29 Electrochemistry
McW Unit VI Oxidation-Reduction and Electrochemistry
Oxidation-Reduction Equations
Electrolysis
Faraday's Laws
The Meaning of a Faraday
Galvanic Cells
DG Set TC-5 Electrochemical and Electrolytic Cells: An Electrochemical Cell (Zn-Cu);
DG Set TC-5 Standard Oxidation Potentials: Hydrogen Electrode;
DG Set TC-5 Standard Oxidation Potential (Cu-H_3O^+);
DG Set TC-5 Electrolytic Cells;
DG Set TC-5 The Lead Storage Cell.

Programmed Instruction:

Barrow, G. M., *Understanding Chemistry, III, Chemical Reactions*, Benjamin, 1967, pp. 41–70, Oxidation Numbers, pp. 71–114, Balancing redox equations.
Kass, G. A., *Individual Study Program in Chemistry, Volume VI Redox and Electrochemistry*, Benjamin, 1969.

CHAPTER NOTES CHAPTER 22

Chapter Note: Activities and Demonstrations: Many of the demonstrations and other instructional aids may have been used in previous chapters. Their reuse here does not necessarily detract from their instructional utility since they motivate and do maintain student interest. The chapter focuses on variations within a period and requires the processes of application, organization, and consolidation of concepts and skills.

Section 22.1. **Laboratory Experiment 44.** "Elements of Period Three." Cut sodium with a knife, scratch magnesium, aluminum, and silicon. Attempt to make a mark on paper with the silicon (recall graphite). Cut white phosphorus under water. Shred a piece of white phosphorus in air. Crush roll sulfur. Use a Tesla coil to activate an Argon lamp.

Section 22.2. **Exhibit models:** crystal sodium (body-centered cubic), magnesium (hexagonal close packing), aluminum (face-centered cubic close packing) to explain melting and boiling point variations. Refer students to *Chemistry Problems*, Metallic Crystals, pp. 372–382, and later to Nonmetallic Crystals, pp. 385–386. Exhibit silicon crystal models (Klinger KC8060, KC8061, KC8062) or use diamond model HRW CH-7154 as representative.

Section 22.3. **Activities and transparencies:** List the eight elements with their atomic numbers in a vertical tabular form. Have students write the electronic configurations. Add the first ionization energies. Have the students explain the variations in ionization energies as they relate to the electron configurations. You may wish to use the orbital configurations as well. Review details of atomic radii (Sec. 5.5) and atomic radii versus ionic radii (Sec. 6.8).

Section 22.4. **Demonstrations:** Burn Na (Caution) and Mg in air. Add water to the products, using Na_2O_2 if desired. Test resultant solutions with litmus. It may be necessary to heat the MgO with H_2O and use phenolphthalein, or pH paper. You may wish to use the equation $O^{--} + H_2O \rightarrow 2OH^-$ for oxide ion reaction with H_2O. Dilute and filter a "Milk of Magnesia" suspension and test the filtrate for alkaline reaction.

Demonstrate the amphiprotic nature of $Al(OH)_3$, preparing it and testing according to Procedure 5, Exp. 54, "Aluminum and Its Compounds." Procedure 1 of this experiment describes the production of Al_2O_3 by amalgamating Al foil with H_g (I) nitrate. For a variation of this demonstration, see Alyea p. 205, "Aluminum: Oxidation, Reduction," in which amalgamated Al reacts with H_2O liberating H_2. Also see Alyea, p. 142 "Burning Magnesium in Steam."

Don't overlook the fact that aluminum is protected by a coating of Al_2O_3. It is a fascinating experiment to remove this coating and see the reaction of aluminum with the oxygen in the air. Run a sheet of aluminum foil back and forth through a solution of mercury(II) chloride. A deposit of mercury forms on the aluminum, although it may be invisible. Hang up the treated foil on a clothesline. Within a few minutes crystals of aluminum oxide form. There is evidence of heat. Catch the mercury(II) chloride that drips from the sheet. Dispose of it properly. Use gloves and wash them too.

Section 22.5. **Demonstrations:** Prepare H_2S from FeS and HCl and test its water solution with litmus. Burn a hydrogen jet in a bottle of Cl_2 (Procedure 4, Exp. 61). Insert moist litmus paper into the product after noting change in color of the Cl_2. Also blow breath across the bottle mouth after HCl has been produced. Prepare and collect a bottle of HCl(g) by displacement of air using NaCl and H_2SO_4 as reagents. Add water and test with litmus. Test for the relative strengths of H_2S and HCl using pH paper. See Alyea, p. 172, "Reactions of metals with phosphorus for production of magnesium phosphide, Mg_3P_2 and of phosphine PH_3." Other phosphine demonstrations are in Fowles, pp. 296–298, and silane demonstrations in Fowles, pp. 287, 288. Write equations representative of Sec. 22.5 as demonstrations are performed. **Question and table:** Question 18(b); Table 22-4.

Films:

AIM Short Yale films (16mm)
YF-224 *Melting Points—Determination and Trends* (9 min, c, s)
YF-255 *Melting Points (Relation to Periodic Table—Summary)* (3 min, c, s)
YF-247 *Metals and Non-metals* (9 min, c, s)
YF-249 *Sulfur, Its Physical States and Properties* (9 min, c and s)
YF-213 *Sodium Peroxide and Water* (2 min, c, s)

YF-232 *Solutions of Alkali Metals in Liquid Ammonia* (Abridged, 5 min, c, s)
YF-221 *Phosphine Smoke Rings* (2 min, c, s)
MLA CHEM 4112 *Chemical Families* (16mm, 22 min, c, s)
CHEM 4151 *Ionization Energy* (16mm, c, s)
CORF *Sulfur and Its Compounds* (16mm, 13+ min, b-w, c, s)
Sulfur and Hydrogen Sulfide (16mm, b-w, c, s)
Silicon and Its Compounds (16mm, 13+ min, b-w, c, s)
Phosphorus (16mm, 18+ min, b-w, c, s)
Chemical Properties of Water (16mm, 13+ min, b-w, c,s)
SUTH *Chlorine—A Representative Halogen* (16mm, 15 min, c, s)
AIM *Commercial Production of Chlorine* (16mm, 9+ min, c, s)

Loops:
MLA CHEM *Alkali Metal Reactions with Chlorine and with Water* (s-8, c)
Holt-Ealing *Production of Sodium-Electrolysis* (s-8, c)
MH 669027 *Combustion II-Burning Phosphorus* (8, 3-min, c)

Filmstrips:
DG 652 The Chemistry of Period Three (40 frames, c)
DG 684 Hydrides and Oxides (40 frames, c)

Transparencies:
HRW CH-7181 (Fig. 22-1)
CH-7182 (Fig. 22-2)
CH-7154 (Fig. 16-3)
K & E 6120-13 Allotropes of Sulfur,
6120-19 Ionic Compound-Magnesium Oxide,
6120-51 Four Kinds of Elements,
6120-60 Oxidation Numbers,
6120-65 A Nelson-type Cell,
6120-67 Electric Furnace for the Preparation of Phosphorus,
6120-74 Production of Active Metals
HM 2-10526 Trends in Ionization Energies
C T-600-18 Ionization Potentials,
T-600-24, 25 Oxidation States of Elements
McW Electronic Configuration of Elements in Periods 1, 2, 3
Ionization Energies
Physical Properties Related to Bond Type
Properties of Group VIIA Elements
The Metallic Bond
Properties of Group IA and IIA Elements
W Slides The Aluminum Compounds
Aluminum
Preparation and Properties of Halogen
The Ionization Tendencies of Metals
The Allotropes of Phosphorus
The Properties of Sulfur
The Oxide of Sulfur
The Properties of Sulfuric Acid
The Preparation and Properties of Hydrogen Sulfide
Sodium

TEACHER NOTES CHAPTER 23

Chapter Note: Activities and Demonstrations: Some of the demonstrations, activities, and experiments have undoubtedly been used during earlier parts of the course, i.e., reactions of sodium and potassium with water, hydrolysis of salts, spectral lines, etc. The occasion may arise when a repeat of a demonstration is useful.

Section 23.1. Exhibits and activities: samples of metals—Na, K in chunk form with smaller pieces under liquid hydrocarbon. Small sealed ampules of Cs and Rb may be obtained through Kawecki Berylco Industries Inc., 220 E. 42 St., New York, N.Y. 10017. Discuss safe storage as it is related to chemical activity. Review the electron configurations of the atoms and ions in table form. Use HRW transparencies CH-7115 and 7120 (Radii), CH-7117 (Ionization Energies) as sources of data which students add to their table. Add oxidation states and use Table 23-1 as summary. Have students explain the long temperature range of metals, melting point to boiling point as indicative of the strength of the metallic bond.

Section 23.2. Activities: Use CH-7117 transparency as basis of discussion of variation of ionization energies with atom size. Refer to Table 21-3 Relative Strengths of Oxidizing and Reducing Agents, Table 21-4 Standard Oxidizing Potentials, and Table 23-2 to raise the problem of lithium position and its relatively high ionization energy. Explain the problem in terms of the small ion size of lithium(I), the consequent high charge density and hydration energy. Expose samples of NaCl and LiCl on separate watch glasses to the air and have students note the hydration of LiCl. Write the generalized question for the electrode reaction. Demonstrate flame tests by using platinum or nichrome wire. Discuss energy acquisition and emission. A variation, spray solutions from squeeze bottles through the blue burner flame. Another variation of flame tests for class demonstration: Place the carbonates of sodium, potassium, and lithium in watch glasses or evaporating dishes. Add HCl. The burner is placed with the collar openings immediately above the reacting substance, (Care!) This results in a colored burner flame which is not as fleeting as the flame using the wire. **Safety:** Discuss the safety storage of potassium and sodium. Use the article by Lehrfeld, J., "A Convenient and Easily Handled Source of Sodium," *Journal of Chemical Education* 48, 279(1971). It suggests the use of a granular sodium-lead alloy, "dry-Na," a product of J. T. Baker Chemical Co., Phillipsburg, N.J.

Section 23.5. Exhibit and Demonstration: Compounds (Sec. 23.4.) and demonstrate solubility and refer to Solubility Chart, Table 12 in the Appendix.

Historical production of sodium: Alyea, p. 178, "Electrolysis of Solid Sodium Hydroxide; Loop: E 80 3940/1 *Production of Sodium-Electrolysis*; Pearce, W. L., "Humphrey Davy," *Scientific American*, June 1960. *Commercial production:* Fowles, G. *Science Master's Book Series IV, Part 2: Chemistry*, pp. 143–144, "Downs process of Na." **Demonstration:** Melt NaCl in graphite crucible with iron rods as electrodes. Cathode has a glass tube sleeve (3 in.) at the end to be inserted into the melt. This jacketed electrode results in the production of a sodium mirror on the inside of the glass. Cool the hot electrode in salt, crush the glass tube, add pieces to H_2O in test tube, test for hydrogen and for NaOH with indicator. The crucible is mounted inside an insulating metal can, asbestos cement-lined jacket and heated with a high temperature burner. See Alyea, pp. 133, 214 for demonstrations for formations of sodium mirror in an incandescent light bulb. Develop the diagram for the Downs cell (Fig. 23-4) and the equations for the process.

Sections 23.6–23.7. **Demonstration:** The physical properties have been previously demonsdrated and discussed. Discuss reaction with air, Na_2O_2 and Na_2O (include equation for NaOH + Na), demonstrate $Na_2O_2 + H_2O$. Demonstrate burning sodium in chlorine (Alyea pp. 146, 201, 212). CHEM loop: 168W1003(T) or 168W1023(K) "Alkali Metal Reactions with Chlorine and with Water." Discuss the uses of sodium in Sec. 23.7. Explain production of tetraethyl lead from ethyl chloride and a sodium-lead alloy. Display and operate a sodium lamp. Discuss the uses of sodium as a heat transfer agent in exhaust valves of engines and atomic reactors. **Questions:** 1–3, 16, 22.

Section 23.8. **Exhibits:** Display rock salt, table salt, iodized salt, salt tablets, halite crystal. Examine NaCl crystals with a hand lens and exhibit a model of NaCl facecentered cubic structure. Have students report on the Morton Salt Co. booklet, *Salt* (history, methods of production, uses). Mention the location of industrial chemical plants near available sources of NaCl such as in upper New York State near Syracuse. Refer students to Block, M. R., "The Social Influence of Salt," Scientific American, July 1963.

Section 23.9. **Demonstrate:** The electrolysis of brine using a large U-tube, carbon electrodes, and vent for hydrogen collection at the cathode. Insert red and blue litmus papers at each electrode. Test for hydrogen collected in a test tube by air displacement. Identify chlorine by waft test and the bleaching of both litmus papers. Identify NaOH by litmus change. The Hoffman apparatus may be used with carbon electrodes. **Refer:** to Fowles, *Lecture Experiments in Chemistry*, p. 431, and to pp. 125, 126, Exp. 151 for the preparation of NaOH from soda; and to Fowles, *Science Master's Book Series IV-Part 2-Chemistry*, p. 98, NaOH in the Hg-brine cell.

Section 23.10. **Demonstrations:** Pass CO_2 through a wash bottle to remove excess HCl. Then pass it into a 28% solution NH_3(aq) saturated with NaCl. If a Kipp generator is not used, dry ice pellets may be dropped into the NH_3-NaCl solution. Start early to allow time for $NaHCO_3$ to precipitate. During this time exhibit commercial samples of washing and baking soda. Write equations for the Kipp and commercial methods for CO_2 production. Demonstrate the laboratory preparation, collection of NH_3 (Chap. 27 and Exp. 57); write formula and ionic equation (p. 476). Identify the $NaHCO_3$ precipitate by filtering, heating, and identifying the CO_2, and adding HCl to the cooled residual Na_2CO_3 again identifying the CO_2 by limewater test. Use HRW transparency CH-7184 (Fig. 23-7) in writing the Solvay equations. Refer to Alyea, p. 217, and Fowles, *Science Master's Book Series IV—Part 2—Chemistry*, p. 142 for Solvay demonstrations.

Section 23.11. **Demonstrations and Exhibits:** Display as many as possible of the compounds and their commercial varieties of the compounds listed in Table 23-3. Discuss their uses. Develop the H_3PO_4 equation series (Sec. 23.11). See Exp. 31. "Neutralizing Phosphoric Acid" in *Scientific Experiments in Chemistry*. Use pH paper to determine pH of the three sodium phosphates and write equations for their hydrolysis. Discuss the uses of phosphates and the ecological problems. Exhibit commercial packages which list the phosphate content or state "low in phosphates" or "contains no phosphates." **References:** *Chemistry* "Substitutes for Phosphates in Detergents," 43, July–Aug. (1970); "A Convenient Method of Detergent Analysis," 45, Mar. (1972); "Phosphates in Detergents," 45, Sept. (1972). *Journal of Chemical Education*, "Analysis of Phosphate Detergents," 48, 551 (1971). **Questions:** 4–7, 13, 14, 17–21, 23–25.

Sections 23.12–23.15. **Refer to** *Chemistry*, "Discovery of Potassium," 37, Aug. (1964). Demonstrate the flame test for potassium. Use Fig. 23-8, 23-9. **Demonstrate:** Production of KNO_3^-. To 25 ml of H_2O add 12 g $NaNO_3$ and 10 g KCl. Heat to complete solution and evaporate to one-half volume. Let the solution cool. Filter out crystals. Examine with hand lens and compare with NaCl crystals. Evaporate filtrate further until a second crop of crystals forms. Filter and examine these with the lens or under microscope. Exhibit samples of compounds in Table 23-4. Discuss uses. Discuss substitution of sodium compounds for potassium compounds except in chemical fertilizers. Exhibit a fertilizer such as 10-6-4 and discuss meaning of each number. **Questions:** 8–11, 22.

Sections 23.26–23.27. **Demonstration:** Discuss methods of production of rubidium and cesium. Use cesium photo cell to ring a bell (Alyea, p. 27). Redemonstrate the photoelectric effect: A negatively charged zinc plate connected to an electroscope is discharged by an ultra-violet lamp. It would not be discharged if it had a positive charge.

Sections 23.18–23.19. Activities: Exhibit a prism spectroscope and explain structure. Have students come up and view the line spectrum of a tube neon, mercury, sodium. Refer to Fig. 23-12 and the Welch chart 4856A. Students may use the inexpensive Macalaster Hand Spectroscope (tube and diffraction grating) to examine flames of flame tests or the operating tubes. Discuss the origin of spectral lines as in Sec. 23.19.

Section 23.3. Demonstration: Preparation of Li: LiCl is fused by heating with a Fisher burner. A heating jacket may be made of a tin can (a one-lb coffee can is good) which is lined with about one-half inch of asbestos cement mixed with a little plaster of Paris. (This jacket is useful for other purposes.) A small iron crucible (30 ml) may be used with a carbon anode and a nichrome wire cathode. The end of the nichrome wire is made into a small loop before it is inserted into the molten LiCl. The voltage used is 10–12 V. The Li which is liberated collects as a small droplet inside the loop. This may be cooled by quickly putting it into solid LiCl. When the material has cooled, the reaction with H_2O, liberating H_2 and forming LiOH, can be demonstrated. Consult article by S. Geffner, "Preparation of Metallic Lithium," *Journal of Chemical Education*, 38, A419 (1961) and Alyea, p. 193. Discuss the uses of lithium compounds and alloys including the use of the oxide for removal of CO_2 in space vehicles.

Films:

MLA CHEM 4151 *Ionization Energy* (16mm, 22 min, c and s)

MLA CHEM 4112 *Chemical Families* (16mm, 22 min, c and s)

CORF *The Sodium Family* (16mm, 16 min, b-w, c, s)

PPG Commercial Production of Caustic Soda (16mm, 8⁻ min, c, s, free loan)

Loops:

MLA or Ward CHEM Study Film Loops, 168 W 1003 (T), 168 W 1023 (K) *Alkali Metal Reactions with Chlorine and with Water* (4 min, c and s)

E-80-3940/1 (H-E) *Production of Sodium-Electrolysis* (s-8, 4 min, c)

E 80-2025/1 *Absorption Spectra* (s-8, 4⁻ min, c)

E 80-2884/1 *The Photoelectric Effect* (s-8, 4 min, c)

Filmstrips:

DG 659 Metals of Group IA and IIA (40 frames, c)

DG 666 Metals and the Metallic Bond

Transparencies:

HRW CH-7183 Carbon-Lithium bond in Ethyl-Lithium

CH-7184 A flow diagram of the Solvay process

K & E 30-6120-65 A Nelson-type Cell (1 + 4)

30-6120-74 Production of Active Metals (1 + 3)

McW The Metallic Bond

Properties of Group IA and Group IIA Elements

Energy Changes—Solid Metal to Aqueous Ion

Energy Changes—Formation of an Ionic Salt from the Elements

Properties of Period 4 Metals

W Slides: Lithium

Sodium

Potassium

TEACHER NOTES CHAPTER 24

Section 24.1. Activities: Exhibit available samples of metals of Group II and some of Group I. Develop a chalkboard and student notebook table in which atomic numbers and symbols of Group II are listed. From the atomic numbers, have students state electron configuration and oxidation number. As in Chapter 23, as source of data, use HRW transparencies CH-7115 and 7120 (radii) for general trend with increasing atomic number. (Record metallic radii from Table 25-1 with discussion of differences between the transparencies and Table 25-1, if desired.) Use CH-7117 (ionization energies). Refer to Table 21-3, Strengths of Oxidizing and Reducing Agents and Table 21-4, Standard Oxidation Potentials. Record the data you wish on the chalkboard table. Compare radii of atoms and ions (Table 6-3).

Section 24.2. Activities: Exhibit samples of veryl and aquamarine. Have students report on the article, "Beryllium," *Chemistry*, 44, Dec. (1971). Discuss its preparation, properties, and uses. Stress its nonmetallic properties (covalence, amphiprotic hydroxide) and the poisonous nature of the element and its compounds.

Section 24.3. Exhibit and Demonstration: Exhibit dolomite, carnallite, asbestos, talc. Exhibit articles representative of the uses of asbestos. Demonstrate its noncombustibility.

Section 24.4 Discussion and Demonstration: Use the HRW transparency CH-7185 as the basis for discussion and development of the important equations. See Fowles pp. 434–435 for demonstration of preparation of magnesium (and calcium) by electrolysis. It is possible to obtain magnesium by electrolysis of molten $MgCl_2$. However, students who have tackled this as a project have found it almost impossible to obtain the molten $MgCl_2$ unless HCl is carefully and continuously added to the melt which forms initially from the hydrate due to extensive conversion to the hydroxide-oxide by heat. Samples of crystalline magnesium and *Dowmetal* may be obtained from the Dow Chemical Co. Discuss the production of magnesium from magnesium oxide and ferrosilicon.

Section 24.5. Demonstrations: Remove the tarnish from magnesium ribbon with steel wool and discuss it as a self-protective metal (refer to use of magnesium alloys in ladders and outdoor furniture). Burn magnesium ribbon in CO_2 and note the solid deposits (white MgO and black carbon on the sides of the bottle. Write the equation for this combustion.

Section 24.7. Exhibits: Exhibit limestone, calcite,

marble, sea shells, coral, coquina, and chalk. Add HCl to samples and have students note effervescence. Other $CaCO_3$ deposits, i.e., stalagtite, stalagmite, boiler scale, hardened lime mortar may also be similarly tested. Exhibit gypsum. A crystal of selenite may be heated to show that the natural hydrate contains water of hydration. Exhibit plaster of Paris.

Section 24.8. **Reference:** Refer to Fowles, pp. 434–435 for preparation of calcium by electrolysis. Write the equation as well as the one for reduction of CaO by aluminum.

Section 24.9. **Demonstrations:** Demonstrate the action of tarnished calcium turnings with those of a fresh sample with H_2O. Have students note the formation of some solid $Ca(OH)_2$ and test with phenolphthalein. Compare the activity of fresh calcium with that of sodium upon addition to H_2O. Have students "notice" that calcium is above sodium in both Table 21-3 and 21-4. Compare the ionic radii and explain the apparent anomoly in terms of the greater charge of the doubly charged calcium ion. Demonstrate the calcium flame test. Discuss uses of calcium.

Section 24.10. Use exhibits and test for Sec. 24.7. **Films:** *Marble Today and Tomorrow* (Vermont Marble Co.), *The Story of Limestone* (Indiana Limestone Institute of America). **Demonstrate:** The reaction for precipitated chalk. Write the equation.

Section 24.11. **Hard water demonstrations:** Add liquid soap to distilled water, tapwater, and hard water (equation). Simulated natural production: (a) $Ca(HCO_3)_2$—have CO_2 continuously bubbling through a water-filled tall cylinder containing limestone or marble chips. Pass CO_2 through $Ca(OH)_2$ solution until the initial precipitate dissolves; (Procedure 2, Exp. 46). (b) $CaSO_4$—pass CO_2 through another cylinder containing broken chunks of gypsum. Samples from (a) and (b) should be tested with soap solution. Boil samples of both temporary and permanent hardness and retest with soap. Have students note evolution of CO_2 bubbles from heated $Ca(HCO_3)_2$ long before actual boiling starts. Write the equation for softening water of temporary hardness. Demonstrate softening by precipitation of $CaCO_3$ or $MgCO_3$ by addition of Na_2CO_3 to both kinds of hardness samples (equations written). Demonstrate softening by $NH_3(aq)$ or limewater addition to temporary hard water and write equations. Try borax and sodium phosphate as softening agents. Demonstrate softening by zeolite column (Welch 5756). Demonstrate cation-exchange, anion exchange, and mixed bed resin exchangers. See Castka, J. and Crane, J., "Demonstrations of Ion-exchange Resins," *Journal of Chemical Education*, 27, 673 (1950). Details of this article are also described in the N.Y.S. Chemistry Handbook, pp. 127–129.

Section 24.13. **Demonstrations:** Place a small piece of limestone directly on the grid of a high temperature burner. Permit it to cool after heating 15 minutes. Add H_2O to residue in a beaker and test with phenophthalein. To a large chunk of fresh CaO in a beaker, add H_2O. Note evolution of steam and swelling of the chunk to form a mound of powder. Test with indicator. Refer to the reactions and equations of the Solvay process (Sec. 23.10). Relate to this demonstration. Add a chunk of "old" and probably air-slaked lime to H_2O. Add acid (*Caution*) to another chunk of the air-slaked lime.

Section 24.14. **Exhibits:** Exhibit solid $Ca(OH)_2$, whitewash, and limewater. Test with litmus paper, phenolphthalein, and pH paper. Heat a mixture of $Ca(OH)_2$ and NH_4Cl and relate to Solvay process. Test a sample of hardened lime mortar with HCl. Discuss hardening of mortar and write equations.

Section 24.15. **Demonstrations:** See the demonstration on heating selenite (Sec. 24.7). Mix plaster of Paris with water and make a mold or casting. Have students note the rapidity of the initial setting. Write the reversible equation. Exhibit castings, wallboard and posterboard.

Sections 24.16–24.17. **Demonstration:** flame tests, Exp. 45. Exhibit compounds of Ba and Sr, discuss their uses.

Films:

CORF *Hard Water* (16mm, 28 min, b-w, c, s)

Marble Today and Tomorrow (16mm, 20 min, s, c), Vermont Marble Co., Proctor, Vermont 05765

The Story of Limestone (16mm, 20 min, s, c), Indiana Limestone Institute of America, Inc., Stone City National Bank Building, Bedford, Indiana 47421

Virginia Spectacular (16mm, 14-min, c, s), Luray Caverns, Virginia Dept. of Conservation and Development, 911 Broad Street, Richmond, Virginia 23219

Loops:

HR Group IIA Elements: Physical and Chemical Properties (s-8, 4 min, c)

Filmstrips:

DG Metals of Group IA and Group IIA (40 frames, c)

Transparencies:

HRW CH-:185 Diagram of a plant for the production of magnesium from sea water.

K & E 30-61: 0-74 Production of Active Metals (1 + 3)

McW Properties of Group IA and Group IIA Elements

Slides:

The Calcium Compounds

Magnesium

Calcium

TEACHER NOTES CHAPTER 25

Chapter Note: **Activities and Demonstrations:** The reading and demonstration references are arranged to follow the text sequence. The first set deals with generalities applying to the transition subgroups. This is

followed by references to the Iron Family, the Copper Family, and the Zinc Subgroup.

This lengthy chapter affords ample opportunity for a wide choice in emphasis and related activities. Since it deals with most of the world's functional metals, the socio-economic aspects such as depletion of natural resources, competition among nations for available supplies of ore, technological advances, and attempts by society and industry to recognize and solve ecological concommitants may serve as the most relevant aspects of the chapter.

The many instructional aids listed (and this listing is far from a complete one) provide for a variety of student activities which can keep interest at a high level. The large number of experiments also provide for a choice of representative laboratory activities and/or projects. These experiments include: Exp. 47 "Metallurgy of the Transition Metals," Exp. 48, "Tests for Iron (II) and Iron (III) Ions;" Exp. 49, "Oxidation States of Transition Elements;" Exp. 50, "Complex Ions of Copper and Silver;" Exp. 51, "Zinc and Its Compounds;" Exp. 52, "Cobalt Nitrate Tests;" Exp. 53, "Separation of Lead(II), Silver, and Mercury(I) Ions;" Exp. 55, "Borax Bead Tests;" Exp. 56, "Separation of Metallic Ions." Experiments from *Scientific Experiments in Chemistry* are Exp. 47, "Chemical Properties of Metals;" and Exp. 50, "Corrosion of Iron." The list of Western(W) 35mm transparency sets should be used as desired with the particular transition element. Included are the set on "The Lead Compounds" and the set on "The Compounds of Chromium and Manganese." The last set may be used as desired when dealing with the generalities for the Transition Elements. Lead, of course, is not a transition metal but the compounds of lead may be considered in connection with Exp. 53, "Separation of Lead(II), Silver, and Mercury(I) Ions."

Section 25.1. **Exhibit:** samples of period 4 transition elements and representatives of other periods, i.e., silver, mercury, and gold. Have students identify them as metals (observation). Transparency: HRW CH-7116, "Sublevel blocks of the Periodic Table." Activities: Develop a chalkboard and student notebook table (Table 25-1) with symbols and atomic numbers as the given information. Students write the representative number of electrons in sublevels and the electron configuration notation. Refer students to Table 7 in the Appendix, Electronic Arrangement of the Elements, for details of elements beyond period 4. Discuss placement of elements 103, 104, and 105.

Section 25.2. **Activities:** Refer to Table 5-5 or HRW transparency CH-7120 and have students determine that the atoms of the transition elements are small compared to those of metals of Groups I and II. Refer to Table 16 in the Appendix, Properties of Common Elements, and have students determine that the transition metals have melting points that are generally higher than those of other metals, i.e. Ti = 1675°C, Fe = 1535°C, Cr = 1890°C.

Section 25.3. **Experiment and Demonstrations:** Laboratory Experiment 49, "Oxidation States of Transition Elements." Re-demonstrate some quick color changes from Exp. 49 or Alyea pp. 42 (19-17), 45-46 (21-15, 21-21, 103-105, 171, 175, 182). These are indicative of variable oxidation states and the comparative ease of transition (energies of *d* and *s* electrons do not differ greatly). **Demonstration:** Add HCl to steel wool or iron filings to show that iron(II) chloride rather than iron(III) chloride is produced. This reaction indicates that the *s* electrons are generally lost before *d* electron(s). Burn ignited steel wool in chlorine to show production of $FeCl_3$ (higher oxidation state produced with chlorine which is a stronger oxidizing agent than H_3O^+—refer to Table 21-3).

Section 25.4. **Exhibit and Demonstrations:** Exhibit samples of the various colored compounds in Table 25-3 and Fig. 25-2. Point out instances where color is related to oxidation state, i.e. Fe^{++}—Fe^{+++}, and of the anion with which the metal is combined. Discuss color (d electrons) as in Sec. 25.4. Demonstrate white light on opaque surfaces—white, red, blue (complement of light removed). Demonstrate that red light on red surface is reflected and that red light on blue surface makes the surface look black.

Section 25.5. **Laboratory Experiments:** Laboratory Experiment 50 "Complex Ions of Copper and Silver." Re-demonstrate selected procedures from Exp. 40, "Equilibrium and Complex Ions" (Cu and Co) including Instructor's Experiment; Exp. 50 (Cu and Ag).

Section 25.6. **Loop and Demonstrations and Student Project: Paramagnetism:** Loop: E 80 2041/1, "Paramagnetism of Liquid Oxygen" (discuss related electron-dot formula for O_2). Define paramagnetism. **Demonstration references:** Campbell, J. A., "Sealed Tube Experiments" (Exp. 58), *Journal of Chemical Education*, 47, 273 (1970); Alyea: diamagnetism, p. 100; paramagnetism, pp. 208, 215; *Journal of Chemical Education*, 40, A407 (1963); "A Qualitative Test for the Presence of Paramagnetic Species," 41, 450 (1964); "Paramagnetism," 48, 5 (1971). **Student project:** Use a soda-straw balance developed for PSSC physics (as Guoy balance) a strong U-shaped magnet, and small pharmaceutical gelatine capsules containing equivalent minimole(s) of substances to be tested ($MnSO_4$, $FeSO_4$, $CoSO_4$, and $ZnSO_4$). Try MnO_2, $KMnO_4$, and Mn. Develop schematic for manganese similar to that for titanium in Sec. 25.3.

Section 25.7. **Discussion:** Review the similarities in structure and properties based on preceding activities and Sec. 25.2. Discuss the similarities within each subgroup and those across the 6th, 7th, and 8th subgroup.

Section 25.8. **Activities:** Exhibit samples of the three metals and their alloys (Alnico, stainless steel, permalloy, etc.). Develop a chalkboard and student notebook table (Table 25-6) for electron configuration and oxidation number. Have students supply other data

from Table 16, Appendix, Properties of Common Elements. Then develop the equivalent of Table 25-7.

Section 25.9. Discussion: Discuss the abundance of iron. Exhibit samples of hematite, magnetite, taconite, and other iron-containing minerals (ores). Discuss the reasons for the use of taconite and the details of its preparation for use in the blast furnace (Fig. 25-5). Samples of pulverized taconite and the sintered pellets are tested with a magnet. Exhibit iron pyrites. Discuss why it is not used as an iron ore. Develop: the equations in Sec. 25.10 as the structural and functional aspects of the blast furnace and its operation are considered.

Section 25.11. Discussion, exhibits, films, demonstration: Emphasize steel production as a process in which impurities in blast furnace iron are removed by oxidation, the sources of the oxygen (gas or iron ore) developing the equations for oxidative reactions. Discuss use of scrap in recycling (Fig. 25-8). Develop equations for slag formation. Exhibit samples of metals used in alloys and of the alloys.

Section 25.12. Vacuum-degassing: Explain reasons for gas removal. Use Fig. 25-13. Demonstration: tempering, Alyea, p. 104 (21012s).

Section 25.14. Demonstrations, exhibits, references: Demonstrations: steel wool, nails, iron filings + (a) HCl, (b) NaOH. *Rusting:* Moistened steel wool in bottom of test tube (one-hole stopper, long glass tube) inverted in beaker of colored H_2O. *Corrosion:* Alyea, pp. 46, 51, 89, 104, 125, 162, 171; Fowles pp. 73-79; *N.Y.S. Chemistry Handbook* pp. 301-307 (p. 306); *Journal of Chemical Education* 47, A387 (1970). Fig. 25-14, Fig. 25-15. Test Fe_2O_3 and Fe_3O_4 for magnetic properties. Exhibit oxides and the commercial pigments.

Sections 25.15-25.17. Laboratory Experiment 48, "Tests for Iron(II) and Iron (III) Ions." Re-demonstrate Exp. 48 as needed. Demonstrations: Hydroxides—$FeSO_4(aq)$ or $FeCl_2(aq) + NH_3(aq) \rightarrow Fe(OH)_2(s)$ white changing to blue-green. $FeCl_3(aq) + NH_3(aq) \rightarrow Fe(OH)_3(s)$ red-brown. Shake the $Fe(OH)_2(s)$ so that the test tube contents moisten the upper part of the test tube and the $Fe(OH)_2$ adhering rapidly changes to the red-brown of $Fe(OH)_3$. This change illustrates the change of $Fe(OH)_2$ to $Fe(OH)_3$ in the process of electrolytic corrosion. Aqueous sodium hydroxide produces the same results and addition of excess NaOH shows that the hydroxides are not amphiprotic. *Colloidal Iron(III) hydroxide:* To 200 ml of boiling distilled water add 1 ml of 1*M* $FeCl_3$ dropwise to develop the deep red colloid. Compare the color with the $FeCl_3$ solution color. *Hydrolysis*: Test solution of $FeCl_3$ with litmus paper and write the text equations. *Thiocyanate:* Add a solution of $AgNO_3$ to the blood-red $FeNCS^{++}$ solution. The color disappears due to formation of AgNCS complex which has a smaller K_i than the $FeNCS^{++}$. Develop the equations in Sec. 25.17 to complement Exp. 48.

Section 25.18. Laboratory Experiment 52, "Cobalt Nitrate Tests." If necessary re-demonstrate Procedures 2, 3, 4 of Exp. 40, "Equilibrium and Complex Ions." Exhibit a cobalt chloride hygrometer. Heat a sample of $CoCl_2 \cdot 6H_2O$ until the blue anhydrous form results or heat a piece of filter paper soaked in the solution of $CoCl_2$. Exhibit: Cobalt, its ores if available, and its alloys. Exhibit blue cobalt glass and demonstrate its use in Exp. 45 "Flame Tests." Student report: Morral, J., "From Cabul to Cobalt," *Journal of Chemical Education* 34, 185 (1957). Discuss use of cobalt-60 in cancer treatment. Refer interested students to *Journal of Chemical Education* "Synthesis and Reactions of Cobalt Complexes," 46, 508 (1969).

Section 25.21. Exhibits: Nickel, coins (Fig. 25-17), and nickel alloys, nichrome coils (toaster). Add HCl to nickel and compare reactivity with that of iron. Exhibit nickel compounds. Precipitate NiS from a solution of a salt. Discuss the Edison storage battery. Exhibit stainless steel cookware.

Section 25.23. Activities and Exhibits. Develop the chalkboard and student notebook table duplicate of Table 25-8 in the same manner used for the iron family table (Table 25-6). Have students compare details of Table 25-8 with those of Table 25-1. Compare Table 21-3 and Table 21-4, noting the effects of the 18-1 outermost shells of electrons on oxidation states and oxidation potentials. Discuss details for the copper family as in Sec. 25.23. Exhibit samples of native copper, silver, and gold-bearing ore (if available). Gold leaf may be exhibited now or later. Reference: "The Color of Metals," *Journal of Chemical Education*, 48, 270 (1971).

Sections 25.24-25.25. Exhibit: available ores. Laboratory Experiment 47. "Metallurgy of the Transition Metals." Repeat procedures from Exp. 47 as required. Demonstration: Flotation as in Exp. 51 of Alyea, p. 25. See Fig. 25-18.

Sections 25.26-25.27 Exhibit: Copper wire, tubing, sheet and electrical equipment. This establishes its physical properties and its uses in the home in plumbing, heating, gutters, and leaders. Display alloys: brass devices, bronze statue, etc. Direct student attention to condition of weathered bronze statues in park and copper-roofed buildings to indicate corrosion. References: Alyea, p. 158 (decomposition of cupric ion by white phosphorus), p. 165 (Synthesis of Cu(I) sulfide); *Journal of Chemical Education*, "Determination of Copper Colorimetrically," 47, A484 (1970), Wilhelm, D. L., "The Law of Definite Proportion," *ibid.*, 50, 436 (1973) uses the synthesis of copper(I) iodide as a laboratory experiment.

Sections 25.28-25.30. Exhibit: Native silver and silver ores. Relate the recovery of silver to the electrolytic refining of copper. Demonstrate: The replacement of silver by copper. Refer to student results in Exp. 50 or re-demonstrate Procedure 2. Discuss the Parkes process. Refer to Alyea p. 28 for photography demon-

strations and p. 193 for silver plating from an iodide complex.

Sections 25.31–25.32. Demonstrate: that gold leaf mounted between two glass plates is yellow by reflected light. It looks green by transmitted light which indicates thinness and malleability. Test gold samples for copper by addition of a drop of conc. HNO_3 and discuss purity in terms of carats. Refer students: to Hausman, E. A., "The Uses of Nobility," *Chemistry*, 37, Oct. (1964).

Section 25.33. Activities: Develop a chalkboard and student notebook table (Table 25-8) in the same way as similar tables were developed in this chapter. Exhibit samples of zinc, cadmium, mercury. Discuss the relationships. Use HRW transparency CH-7190 (Fig. 25-25) and show the structure of the Hg_2^{++} ion as $[Hg{:}Hg]^{++}$. Compare the data in Table 25-8 with that in Table 24-1 for the Group II metals.

Sections 25.34–25.35. Demonstrations and Exhibits: The metallurgical processes of roasting and reduction should be discussed as being widely applicable for sulfides and oxides. Refer to Exp. 47, "Metallurgy of the Transition Metals," for demonstration procedures. Instructor Experiment—roasting ZnS (of PbS). Use equations in Sec. 25.34. Demonstrate electrolytic process Sec. 25.34-2. Exhibit the various forms of Zn. Laboratory Experiment 51, "Zinc and Its Compounds;" Experiment 56, "Separation of Metallic Ions"; Experiment 52, "Cobalt Nitrate Tests." Demonstration: galvanizing, Instructor Experiment, Exp. 51.

Section 25.36. Exhibit: The metal. Use Fig. 25-24. Precipitate CdS from a salt solution. Discuss uses of cadmium particularly its use in control rods for nuclear reactors.

Sections 25.37–25.38. Exhibit: Hg and cinnabar. Relate metallurgy to the thermal decomposition of HgO. References: *National Geographic*, "Mercury, Man's Deadly Servant," 142, 528 (1972). *Chemistry*, "Mercury Poisoning," 41, Apr. (1968); "Mercury in Fish," 44, Feb. (1971); "New Focus on Mercury," 43, Nov. (1970). A demonstration of amalgamation of Al article, "How to Grow Hair on a Billiard Ball," 42, Apr. (1969). Demonstrations: Hg + acids, amalgamation, Alyea pp. 123, 36, 138, 143, 180, 185.

Filmstrips:

DG 549, "Changing Ores into Metals"
DG 663, "The Transition Elements."

Films:

MLA CHEM-4172 *Vanadium—A Transition Element* (16mm, 22 min, c and s)
CHEM-4072 *Teacher Training Introduction to "Vanadium—A Transition Element"* (16mm, 7 min, b-w, s)
CORF *Chromium and Manganese* (16mm, 37 min, b-w, c, s)
BFI (Bailey Films) *Copper Mining* (16mm, 14 min, c and s)
USS United States Steel Corp., Educational Services, 600 Grant St., Pittsburgh, Pa. 15230. Catalogue of available films and addresses of local distribution centers is available.
The Story of Nickel (16 mm, 30 min, b-w, s)

Loops:

H-E E 80-2041 1 *Paramegnetism of Liquid Oxygen* (s-8, 3^+ min, c)
MLA CHEM Study Film Loops 168 W 1008(T) or 168 W 1028(K) *Ion Exchange* (s-8, 4 min, c)
MH 669028 *The Rusting of Iron* (8mm, 2^+ min, c)
LBF 11 *The Mercury Beating Heart* (8mm, 4 min, c)
E 80-2033 1 *Ferromagnetic Domain Wall Motion* (s-8, (s-8, 4^+ min, c and s) Also available from Macalaster Scientific Corp. (16 mm, 4 min, c and s)
HR *Group IB Elements: Copper* (s-8, 4 min, c)

Filmstrips:

DG 549 *Changing Ores into Metals* (40 frames, c)
DG 663 *The Transition Elements* (40 frames, c)
American Iron and Steel Institute, 150 West 42nd St., New York, N.Y. 10017 (One copy free to a school system). Includes the chemistry of iron, raw materials, and steel.

Transparencies:

HRW CH-7186 A possible space model of a complex model of a complex ion
CH-7187 Magnetism in matter stems from the spin of electrons
CH-7188 A sectional view of a blast furnace
CH-7189 New Technology in the Copper Refining Process
CH-7190 Models of Mercury(I) Chloride and Mercury(II) Chloride
CH-7116 Sublevel blocks of Periodic Table
K & E 30 6120-25 Diagram of *p* and *d* Orbitals (6)
6120-26 Periodic Table and Atomic Orbitals
Slides W The Silver Compounds
The Mercury Compounds
The Copper Compounds
The Iron Compounds
The Compounds of Chromium and Manganese
The Cobalt Compounds
The Lead Compounds

TEACHER NOTES CHAPTER 26

Chapter Note: Activities and Demonstrations: The importance of aluminum as an industrial metal and the importance of solid state devices based on semiconductors and transistors should highlight the activities and demonstration program. There are many opportunities for student reports based on such articles as Garrett, A. B., "The Flash of Genius-Discovery of the Transistor," W. Chockley, J. Bardeen and W. Brattain, *Journal of Chemical Education*, 40, 302 (1963) and "Synthetic Borane Chemistry—Challenge to Chemical Theory,"

Chemistry, 41, Nov. (1968), reprint 81, Many references indicate how students may do related projects, i.e. Meyers, S. S., *Experiments with Conductors and Semi-conductors*, Bell Telephone Laboratories, 1965. For film-oriented students there are many films recommended. The most recent film is the Dow Corning film, *Wondering About Things*. Another stimulating film is *PPG: Glass for the Seventies*. Related laboratory experiments are Exp. 54 "Aluminum and Its Compounds"; Exp. 55. "Borax Bead Tests"; Exp. 56 "Separation of Metallic Ions," and parts of Exp. 44 "Elements of Period Three." Review the details of Chapter 22, Elements of Period Three, which deals with aluminum and silicon.

Section 26.1. **Exhibits and Activities:** Exhibit aluminum and the metalloids. Define metalloid. Transparency HRW CH-7116—mark in the zig-zag and identify the metalloids (Fig. 26-1). Develop a chalkboard and student notebook table with symbols and atomic numbers given. Have students state and write in the electron configurations which may be developed in greater detail for the valence electrons i.e. Al: 2, 8, $3s^2 3p^1$. **Transparencies:** HRW CH-7117 (ionization energies), CH-7115 and CH-7120 (atomic radii) as source of these data to be added to the table. *Refer students:* to Table 16, Appendix, Properties of Common Elements, for density, melting point, boiling point data (and oxidation number(s) if you wish) for many of the elements in Table 26-1.

Section 26.2. Identify aluminum as a metal by its physical properties exhibiting foil, wire, pots, etc. Review details about its oxide, hydride, amphiprotic hydroxide in Chap. 22. Establish the nonmetallic and metalloidal behavior of boron by reference to its small atomic size (Table 26-1), covalence of its compounds, high ionization energy (compared to aluminum), high electronegativity, and increasing conductivity with increasing temperature (Sec. 26.11).

Section 26.3. **Discuss:** Review the history of aluminum. Students are generally impressed by the "coincidence" of the Hall-Heroult discovery.

Section 26.4. **Slides, charts and exhibits:** A set of charts and a materials exhibit (bauxite, refined alumina, clay, cryolite, carbon electrode) may be obtained from the Aluminum Company of America. **Activities:** Develop a chalkboard outline diagram of the electrolytic cell and write the electrode equations. Refer to Fig. 26-3. Discuss the new chlorination methods (Sec. 26.4).

Section 26.5. **Laboratory Experiment 58:** Aluminum and Its Compounds. **Demonstrations:** Al + NaOH; hydrolysis: test $Al_2(SO_4)_3$ with litmus. Mix $NaHCO_2$ with $Al_2(SO_4)_3$ and add H_2O to the mixture of solids, test for CO_2. Exhibit can of "Drano," have students read the label and detect many pieces of Al in the solid mix. Remove the Al from a small sample (*Caution*) and test the residue with litmus to show OH^- ion. Add a sample of the "Drano" to H_2O and note the vigorous gas production but a negative test for hydrogen. Refer to Breedlove, C. H., "Chemical Basis for a Common Type of Drain Cleaner," *Chemistry*, 44, Sept. (1971) for complete description of reactants and reactions. Students expected hydrogen because of Al + NaOH reaction. "Drano" contains $NaNO_3$ and the net equation is:

$$8Al(s) + 5OH^-(aq) + 2H_2O(\ell) + 3NO_3^-(aq) \rightarrow 3NH_3(g) + 8AlO_2^-(aq) + \text{heat}$$

The NH_3 odor is masked by the pleasant odor of inert ingredients. Demonstrate electrolytic cleaning of tarnished silverware (Alyea p. 31). Write all equations in Sec. 26.5.

Section 26.7. **Demonstrations and Exhibits:** Oxide uses—exhibit various forms: bauxite, corundum, emery, alundum, synthetic ruby and sapphire. See Fig. 26-6 and Fig. 26-7. Exhibit articles to indicate uses: emery cloth, grinding wheel, alundum laboratory articles, jewels in watches.

Section 26.8. **Demonstrations and Preparations:** Refer to Exp. 54, Procedure 5 for preparation and amphiprotism. Use equations in Sec. 26.8. Demonstrate that the hydroxide does not form NH_3 complex ion by its failure to dissolve in excess $NH_3(aq)$ (Procedures 2, 5a and 5d Laboratory Experiment 56 "Separation of Metallic Ions.") Test solutions of $AlCl_3$ and $Al_2(SO_4)_3$ for hydrolysis with indicator. Precipitate $Al(OH)_3$ by addition of solutions of Na_2CO_3 and Na_2S or $(NH_4)_2S$. Demonstrate the use of $Al(OH)_3$ as a coagulum and mordant (lake formation using logwood or litmus solution). See Alyea pp. 31 (14-7).

Section 26.9. Exhibit samples of alum including chrome alum. Develop general formula from specific formulas. See Alyea p. 31 (14-13) for growing alum crystals. Test an alum solution for hydrolysis and precipitate $Al(OH)_3$ from an alum solution.

Section 26.10. Exhibit samples—mica, feldspars, clays, Fuller's earth. Demonstrate use of Fuller's earth or a commercial dry spot remover in oil spot removal. Use Western slides—"Aluminum Compounds." Question 1-6, 17-21, problems 1-3.

Section 26.11. **List data:** For electronic configuration, electronegativity, atomic radius, and ionization energy. **Discuss:** Covalent bonding in its compounds and increased conductivity with increased temperature.

Section 26.13. **Discuss:** Norbide and borazon (BN) production and uses. Demonstrations: preparation of boric acid (Alyea p. 33, 15-2); flame tests, Fig. 26-10 (Alyea p. 33, 15-4); boric acid is a weak acid (pH and conductivity); hydrolysis of borax; see Alyea p. 95 (15s), Alyea, *TOPS*, topic 15.

Laboratory Experiment 55: "Borax Bead Tests."

Section 26.15. **Activities:** List the data for silicon as was done for boron which establishes it as a metalloid. Refer to interchangeability of silicon and boron

in glass (compare atomic radii). Demonstrations: Alyea p. 95, 14-3s silicon by thermite; *N.Y.S. Chemistry Handbook*, p. 185 reaction of silicon and steam. References: Holden, A., *Conductors and Semiconductors, Experiments with Conductors and Semiconductors.*

Section 26.16. Films: CORF, *Silicon and Its Compounds*; DOW, *Wondering About Things*; INDU, *Sand and Imagination*. Discuss: the occurrence of silicon and its role in mineral chemistry. Exhibit: ferrosilicon, transistors, semiconductor and "solid state" devices, sodium silicate, forms of SiO_2 (Sec. 26.16), Sic (paper, grinding wheels, sharpening stones).

Section 26.17. Demonstrations: Burn pieces of Silastic and natural rubber. Collecting solid products of combustion on outside of a test tube. "Bouncing putty"—bounce a ball of it, hit a ball of it with a hammer. Pull the ball apart. Immerse filter paper in a mixture composed of five parts of Dow Corning 104 Emulsion, one part of Dow Corning Catalyst No. XF21 and four parts of water. Air dry or dry the paper in an oven. Shake a mixture of toluene with water colored with red ink. Pour this mixture through untreated and treated filter paper. The water goes through the untreated paper and the toluene through the treated paper (indicates water repellancy). See Alyea p. 34, 15-26, 15-27. Exhibit: commercial silicone products. Diagram the silicone structure, Sec. 26.

Sections 26.19-26.20. Student Reports: on the article by Sabine, D., "Arsenic," *Chemistry*, 42, Apr. (1969). Exhibit and Demonstrate: arsenic, lead shot, and $AsCl_3$. Dissolve $AsCl_3$ in HCl and demonstrate hydrolysis on addition to water. Precipitate As_2S_3 from the acid solution of $AsCl_3$ with H_2S.

Sections 26.21-26.22. Exhibit: chunk antimony and crush in a mortar. Exhibit type, metal, anti-friction alloys, storage battery plates. Re-demonstrate dropping powdered antimony into chlorine. Demonstrate hydrolysis of acidified $SbCl_3$. Discuss recovery, uses, and compounds. Demonstrations: Alyea, p. 40 (18-20 through 18-25, p. 100, 18-14s). Questions and Problems: questions 15, 16; problems 7, 8.

Films:

CORF *Silicon and Its Compounds* (16mm, 14⁻ min, b-w, c, s)

INDU *Sand and Imagination* (16mm, 29 min, b-w, s)

LBF *Corrosion III—Aluminum* (8mm, 4 min, c, s)

MLA *Unfinished Rainbows* (16mm, 36 min, c and s)

Borax from Desert to Farm, American Potash Institute, 1102 Sixteenth St., N.W., Washington, D.C. 20006.

What's a Silicone? (32 min, s.c), 1953, Dow Corning Corp., Midland Mich.

Rubber from Rock (28 min, s, c), Dow Corning Corp. 1955.

Silicones (10 min, s, b/w), 1950, General Electric Co., 570 Lexington Ave., N.Y., N.Y.

Glass, Science, and People (20 min, s, c), Owens-Illinois Glass Co., P.O. Box 1035, Toledo, Ohio.

Glass Center of Corning (25 min, s, c), 1953, Association Films, 1180 Jackson St., Dallas, Texas.

Glass and You (28 min, s, c), Association Films, 1955.

Quartz, Crystal Growing (16 min, 16mm, s, b/w), Sterling Movies, U.S.A., Inc., 45 W. 61 St., N.Y., N.Y. 10022

The Wide World of Borax (11 min, 16mm, s and c), U.S. Borax and Chemical Corp., 630 Shatto Place, Los Angeles, Ca. 90005

Alumina Jamaica (20 min, 16mm and 35mm, s and c), Aluminum Limited Sales Inc., Film Dept. for U.S. Distribution, 111 W. 50th St., N.Y., N.Y. 10020.

Aluminum on the March (28 min, 16mm, s and c), Reynolds Metals Co., Motion Picture Dept., 5004 West Clay St., Richmond, Va.

Wondering About Things (16mm, c and s), Dow Corning Corp., 401 Sylvan Ave., Englewood, N.J.

PPG: Glass for the Seventies (16mm, 12 min, c, s), Modern Talking Picture Service, 2323 New Hyde Park Road, New Hyde Park, N.Y. 11040, or local office, free.

MLA YF-229 *Hydrogen Generator— March Test* (16mm, 7 min, b-w)

Filmstrips:

DG 650 *The Chemistry of Carbon and Silicon* (40 f)

DG 668 Boron and Aluminum (40 f, c)

Transparencies:

W slides The Aluminum Compounds
Aluminum

TEACHER NOTES CHAPTER 27

Chapter Note: Activities and Procedures: The central theme of this chapter is the Haber process, discussed as a contributory cause of World War I. Most students are acquainted with the names of conventional explosives and, from such names, realize that nitric acid is necessary to their manufacture. The process was worked out commercially by 1913 and in 1914, Germany took the initiative in starting a World War. The significance of Germany's previous dependence on Chilean nitrate for fertilizers and nitric acid and the Allied control of the sea may be discussed. This may involve consideration of natural resources, stockpiling, and the effect of chemical discovery on a changing world economic, industrial, and political situation. This type of approach leads to appreciation of the importance of nitrogen as a source of food supply and explosives; and to the problems of the air as a source of nitrogen; and the properties of oxygen and nitrogen fixation. Related ecological problems involving nitrogen compounds include water supply and pollution. The steady addition of ammoniacal wastes and the excessive growth of blue green algae as well as the presence of excess nitrates in water supplies focus on problems

of sewage disposal and water purification processes. These problems should be considered at appropriate times during the chapter development. Student experiments include: Experiment 57, "Ammonia, the Ammonium Ion, and the Nitrate Ion" and Experiment 14, "Percentage of Oxygen in Air," which may have been performed previously. Experiments from *Scientific Experiments in Chemistry* are: Experiment 53, "Properties of Nonmetals:" Experiment 54, "Ammonia;" Experiment 55, "Preparing Nitric Acid and Finding Its Concentration;" Experiment 49, "Reaction between Zinc and Nitric Acid."

Sections 27.1-27.3. Demonstrations: Prepare a sample of oxygen-free air by burning phosphorus in a sample of air (confined in a graduated cylinder). Test residual gas with flaming splint. After shaking liquid contents to dissolve P_4O_{10} test with litmus. See Laboratory Experiment 14, "Percentage of Oxygen in Air," which students may perform. Oxygen-free air may also be prepared by passing air over heated copper or by the oxidation of moistened steel wool or iron filings in a slow oxidation experiment set up the day before, Alyea p. 7 (2-13).

Section 27.4-27.7. Discuss: Physical properties. For solubility refer to Fig. 27-2 and mention "bends." Demonstration: expose white phosphorus to the air, burn red phosphorus in a deflagrating spoon. Discuss nitrogen triple bond, the bond energy, and chemical inactivity. Review the bonding in P_4 (Fig. 22-1). Discuss relative activities of O_2 versus S_8, F_2 versus Cl_2 and N_2 versus P_4. Refer to data on bond energies in Chap. 6. Test for nitrogen and its presence in air: Burn a wad of magnesium ribbon in a crucible. Add a drop or two of water to the hot mix. Test for odor of NH_3 by odor and moistened red litmus paper held over the crucible as the H_2O is added. Write equations (Sec. 27.4). Mention that the burning splint test is nonspecific for nitrogen since other gases (CO_2, SO_2, etc.) do not support this combustion. Discuss the uses of nitrogen as in Sec. 27.6. Display a "burned-out" tungsten filament bulb calling attention to the gray deposit.

Section 27.8. Demonstration: Pass sparks from an induction coil or Tesla coil through a sample of confined air. The electrodes may be of copper inserted through stoppers into a glass tube. A piece of inserted moistened blue litmus turns red. On occasion some color change to nitrogen dioxide may be observed. See Alyea p. 38, 17-10, p. 99, 7-17s. Exhibit: Roots of legume having nodules of nitrogen-fixing bacteria. Refer to *Foundations of Chemistry*, p. 142 and Fig. 4-22, Nitrogen Cycle. Write equations for $CaCN_2$ (Sec. 27.8) and display a fertilizer example. Mention possible use of nuclear reactors.

Sections 27.9-27.10. Laboratory Experiment 57. "Ammonia, Ammonium Ion, and the Nitrate Ion." Redemonstrate procedures in Experiment 57 as required. Write the equations in Sec. 27.10. Demonstration: Destructive distillation of bituminous coal. The apparatus consists of a test tube connected to a collecting bottle. The delivery tube leads almost to the bottom of the bottle which also has a short glass jet exit tube. A large piece of moistened red litmus paper is inserted into the bottle and held in position by the stopper. Heat the coal. (Canned coal works very well.) The coal gas is ignited at the jet exit. The litmus turns blue. Write the equation for the $NH_3(g) + H_2SO_4(aq)$ reaction. *Haber process:* HRW transparency CH-7191 (Fig. 27-5). Discuss the conditions under which the process operates in the light of application of the Le Chatelier principle and equilibrium. Simultaneously demonstrate: The catalyst speeds up the reaction in both directions. Show the effect of high temperature on the equilibrium by passing dry NH_3 (prepared by $NH_3(aq)$ dropping onto pellets of NaOH in generator). Pass NH_3(dry) over heated steel wool. Collect the products by displacement of water. Test for hydrogen.

Sections 27.11-27.12 Demonstrate: The ammonia fountain (Alyea p. 12, 27, 149, 174) or use film: YF-220 *Ammonia Fountain.* Use these and invert a bottle of stoppered $NH_3(g)$ into water (test with indicator) to show great solubility of NH_3. Use models and the chalkboard electron-dot equations to illustrate the $NH_3 + H_2O$ reaction. Demonstrate: liquefaction of NH_3. Dry NH_3 (prepared as above) is passed into a test tube cooled by DRY ICE acetone mix. See Alyea p. 37, 17-6 for burning NH_3 in O_2. If desired redemonstrate conductivity of NH_3^+ toluene. Add H_2O, mix and retest for conductivity. Bring moistened stoppers from $NH_3(aq)$ and conc. HCl together. Write the equation and electron-dot equation. Discuss the $NH_3 + HCl$ reaction as a test for NH_3 liberation of NH_3 in test for $NH_4^+ + OH^-$). Differentiate between NH_3 and NH_4^+ using electron-dot formulas and models.

Section 27.13. Ammonia uses: Activities: Display a bottle of household ammonia. Have students read the label, detect the odor, and test the vapor with litmus paper. Refer to TV commercials mentioning ammonia. Demonstrate cleaning glass with $NH_3(aq)$ versus cleaning with NaOH solution. Have students test for the odor of rubber latex, then coagulate with $HC_2H_3O_2$. Discuss the uses of ammonia and ammonium compounds as in Sec. 27.13. Refer to Fig. 27-9, Fig. 27-10, Fig. 27-11.

Section 27.15. Demonstrations: 1. *Laboratory preparation from $NaNO_3 + H_2SO_4$ (Fig. 27-11.1).* Have students attempt to explain why a retort is used. Why does the acid react with copper? Why is there no evidence of free hydrogen when the acid reacts with zinc? The last two questions are based on an initial demonstration of these facts. The highlight of the lesson is the demonstration of the oxidation reaction of the acid on excelsior and/or the ignition of a glowing splint in the vapors of a heated sample of the acid produced in the retort preparation. (Hold the splint with tongs to prevent staining of hand by the vapors and make sure that the mouth of the test tube is

pointed away from the instructor and the class.) What does the method of collection reveal about HNO_3 (tapwater may be used instead of ice water). Write two forms of the equation. Use Western slide, "Preparation and Properties of Nitric Acid," Questions 12, 13.2. *From ammonia-Ostwald process.* Demonstrate the catalytic oxidation of ammonia to oxides of nitrogen. Dry NH_3, prepared by dropping conc. $NH_3(aq)$ on pellets of NaOH, is passed over heated platinized asbestos with the simultaneous admission of air through a two-hole stopper at the reaction tube inlet. The outlet tube leads into a large flask in which the brown NO_2 becomes visible. This is then dissolved in water and tested with litmus, and for the nitrate ion. To avoid formation of a white cloud of NH_4NO_3, preheat the catalyst before starting NH_3 production. (An electrically heated coil of platinum wire may be substituted for the platinized asbestos.) See Alyea p. 28, 17-11, 17-12; p. 169; p. 179; p. 188; p. 224; and *Journal of Chemical Education*, 41, A973 (1964); 47, A534 (1970). Write the equations in Sec. 27.15. Use HRW transparency CH-7192 (Fig. 27-12). *Related demonstrations are:* (a) Open a flask of NO to air and have class observe change. Add H_2O, stopper and shake the flask. Reopening the flask produces more NO_2. (b) Prepare and collect sample of NO and NO_2. Test solubility, reaction with H_2O, and other properties. Open small bottle of NO. The NO turns brown. Add a small amount of H_2O to the flask. Moisten your palm and use it to cover the bottle opening. Shake. The glass bottle held vertically adheres to the hand. The yellow stain on the demonstrator's palm indicates the formation of xanthoproteic acid (a test for proteins which may be demonstrated by heating boiled egg white with HNO_3).

Sections 27.16–27.18. Activities: Refer to method of collection and discuss details in Sec. 27.16. Heat the sample of HNO_3 collected in the laboratory preparation (*Caution*). Insert a glowing splint into the vapor. Remove splint. Into the mouth of the test tube place a loosely packed wad of wool or excelsior. Reheat the HNO_3. Write the equation in Sec. 27.17 indicating instability, relate to use of retort and HNO_3 as an oxidizing agent. Add HNO_3 to zinc and copper. Test for hydrogen and note production of oxides of nitrogen. (NO_2 observable.) Ask why no hydrogen is produced with zinc and why copper reacts. Write equations for reaction of copper with HNO_3 (dil. and conc.). To show that hydrogen is produced with an active metal use magnesium and HNO_3 (1:25 or greater dilution). Write the equation. Add conc. HNO_3 to gold leaf. Then add conc. HCl. This combination shows the action of aqua regia. Demonstrate the test for a nitrate Sec. 27.18. Discuss the uses of HNO_3 in Sec. 27.19.

Reference: *Foundations of Chemistry* pp: 27–29 Water pollution; pp. 396–399 Photochemical Smog; pp. 606–608 HNO_3; "Our Environment Battles Water Pollution"; A Primer on Water Pollution"; *Chemistry*, Reprint 104, 45, Jan. (1972) "Smog Formation Simplified."

Films:

MLA CHEM 4136 *Nitric Acid* (16 mm, 18 mn, c,s)

CORF *Preparation and properties of Nitric Acid* (16 mm, 21 min, b-w, c,s)

Nitrate, Nitrogen Story (16 mm, 18 min, c.s) California Chemical Co. Ortho Division, 200 Bush St., San Francisco, Cal.

How To Use Nitrogen Solutions (16 mm, 16 min, c.s) Allied Chemical Corp., 40 Rector St., N.Y., N.Y. 10006

AIM Yale Films

YF-220 *Ammonia Fountain* (16 mm, 2 min, c, silent)

YF-232 *Solutions of Alkali Metals in Liquid Ammonia (Abridged)* 16 mm, 5 min, c, s)

YF-218 *Condensation from Air* (16 mm, 1 min, c, s)

CORF *Nitric Acid Compounds and the Nitrogen Cycle* (16 mm 19 min, b-w,c,s)

Nitrogen and Ammonia (16 mm, 16 min, b-w, c,s)

Nitrous Acid and Sodium Nitrite (16mm, 19 min., b-w,c,s)

YF 217 *Distillation from Liquid Air* (16 mm, 2 min, c,s)

Filmstrips:

DG 638 Nitrogen and Its Compounds (40 f, c)

Transparencies:

HRW CH-7191 A flow diagram of the Haber process.
CH-7192 A flow diagram of the Ostwald process.

HM 2-10534 The Nitrogen Molecule
2-10536 The Ammonia Molecule-Hybridized

K&E
30-6120 Series
Covalent Compounds—Nitrogen Molecule (1 + 3)
21 Covalent Compounds—Ammonia (1 + 4)
35 The Mole of Nitrogen (6)

Slides:

W Preparation and Properties of Ammonia
Preparation and Properties of Nitric Acid

Kits:

La Motte Chemical Products Co., Chestertown, Maryland 21620

Water Test Sets: 4795 Ammonia Nitrogen in Water
7485, 7428 Nitrate in Water Test Set
7421 Nitrite in Water Test Set (Model PLN)

Soil Test Sets: 5017 Nitrate in Soil Laboratory Test Set (Model St-N)

General Book References:

See the end of the teacher's Section at the front of the text.

Book references for this chapter:

Simpson, C. H., Chemicals from the Atmosphere, Dept. SIS, American Chemical Society (1973)

Journals:

Chemistry

Vol. 41 (1968) "Liquid Ammonia, A Unique Solvent," Apr. Reprint 78: Molecular Structure, solvent properties, reactions, and metal-ammonia solutions are described.

Vol. 44 (1971) "Pollution: Phosphates and Nitrates," Mar.

Vol. 45 (1972) "Smog Formation Simplified," Jan.

Journal of Chemical Education

Vol. 45 (1968) TOPS: A567, A O A 673-674 (Phosphorus Family)

Vol. 47 (1970) TOPS: p. A534 (Arc process, Formation of NO and NO, Redox reactions involving iodine (HNO, S203).

Vol. 48 (1971) "Gas Laws, Equilibrium, and the Commercial Synthesis of Nitric Acid," p. 838

"Nitrate Analysis for Nonscience Major," p. 647

Vol. 49 (1972) "Reducing Properties of Ammonia," p. A85

Reader's Digest

Sept. (1972) "The 'Clean-Auto' Law is the Wrong Answer," A. Q. Maisel.

TEACHER NOTES CHAPTER 28

Chapter Note: Activities and Demonstrations: Current concern with ecology and pollution indicates that the relation to sulfur and sulfur compounds to air pollution may be made the focal point of the chapter. Interested students should be referred to Weaver, E.C., *Scientific Experiments in Environmental Pollution*, Holt, Rinehart and Winston, 1968: Experiment 4 "Sulfur Dioxide in Air," Experiment 5 "Total Acids in Air (I), "Experiment 6 "Total Acids in Air (II)." Another worthwhile reference is Wartburg, A.F., Lodge, J.P. Jr., "Estimating Concentration of Air Pollutants," *Chemistry*, 41, Feb. (1968) which includes testing for sulfur dioxide. An early student report on sulfuric acid such as Schefield, M., "Early Days of Sulfuric Acid," *Chemistry*, 45, Oct. (1972) may be used to develop the importance of sulfur in industry.

Sections 28.1-28.3. **Discussion and demonstration:** Precede the discussion of the Frasch process by demonstrating (a) insolubility of sulfur in H_2O, (b) the low melting point of sulfur. Use HRW transparency CH-7134 as the basis for a discussion of superheated water. **Demonstration model:** *Science Masters' Book, Series IV-Part 2 Chemistry*, p. 114 uses a modified Liebig condenser and yellow wax. The color and phase changes, Sec. 28.3 are demonstrated in preparing monoclinic and amorphous sulfur.

Section 28.4. **Laboratory Experiment 58.** "Forms of Sulfur." **Demonstrations:** Re-demonstrate as required. If you do not have students use CS_2, demonstrate preparation of rhombic sulfur. Crystals of good size are obtained by covering the watch glass or evaporating dish with a large inverted beaker. It is necessary for a support so that a small opening at the beaker mouth permits slow evaporation of CS_2 (*Caution*: Use hood or window sill). Examine crystals with hand lens. The various changes in color, phase, and viscosity are demonstrated and tabulated with significant temperatures added (Secs. 28.3, 28.4). **Discuss the models:** Exhibit ball-stick model of S_8 (Exp. 44) using carbon balls painted yellow. Open the S_8 model and add more balls for Fig. 28-25.

Section 28.5. **Demonstrations:** (a) Burn a chunk of sulfur in a deflagrating spoon. Insert the spoon into a bottle using a cover plate. Insert a piece of moistened blue litmus paper. Have students detect odor by waft test. Extinguish burning sulfur by submersion in H_2O. (b) Insert a 6-inch strip of copper into the vapors of boiling sulfur. (c) Ignite (*Caution*) a small cone of a mixture of zinc powder and flowers of sulfur on an asbestos sheet. Write equations for (a-c).

Section 28.7. **Discussion:** Mention air pollution but defer details until Sec. 28.15. **Exhibit:** rotten egg and prepared solution of H_2S.

Section 28.8-28.11. **Laboratory Experiment 59,** "Hydrogen Sulfide and Qualitative Analysis." **Demonstrations:** use procedures in Experiment 59 as desired. These cover Sec. 28.8-Sec. 28.11. Develop equations in these sections in conjunction with demonstrations. Other demonstration suggestions: (a) *Introduction:* Good starting demonstration for introducing H_2S and its use in analysis is the precipitation of the colored sulfides. Turn your back and have students pour a sample of each labeled salt solution into separate test tube. You add H_2S solution to each test tube and identify the solution cation. Then add H_2S to a solution of a different metal which results in the same color (*i.e.* black, yellow) as one of those you identified. This poses the problem of procedures for separating and identifying cations by use of H_2S in qualitative analysis. (b) *Burning H_2S:* 1. Have the jet flame of H_2S imping on a cold surface to show production of sulfur. 2. Have the jet burning inside a large beaker containing a dilute solution of $KMnO_4$. (c) *Reducing agent:* 1. H_2S to acidified $KMnO_4$ and/or to $FeCl_3$ solution produces color changes and production of sulfur. 2. Bottles of H_2S and SO_2 mouth to mouth with withdrawal of cover plates. 3. $H_2O_2 + H_2S$. Equations for these reactions provide for review of writing oxidation-reduction equations. (d) *Poisonous nature of H_2S:* 1. Use of hood in preparation and use. 2. After using the H_2S generator dispose of the excess

and stop the reaction by igniting the gas at a jet tip and pouring water through the funnel tube to fill the generator. (e) *Tarnish:* 1. Put a silver plated utensil into $H_2S(aq)$, Equation: $4Ag + 2H_2S + O_2 \rightarrow 2Ag_2S + 2H_2O$. 2. Electrolytic cleaning, Alyea p. 31, 14-15. (f) *Test for H_2S:* Lead acetate paper test. Test for pH of H_2S solution. Develop equations ionization constants for H_2S, and mathematical relationships as in Sec. 28.10-3, and Table 28-2.

Section 28.12. **Demonstrations:** See Procedure 7 in Experiment 59 for demonstrations related to the principles of qualitative analysis. Have students perform the calculations of the sulfide ion concentrations in Table 28-2. Perform the calculations in Sec. 28-12 for the two groups of sulfide precipitations. Use Table 28-3. Use problems 13-16. **Demonstration:** Use of thioacetamide as a convenient source of H_2S in analysis. Heat a solution of thioacetamide which is slightly acidified with HCl. The equation for the hydrolysis of thioacetamide is: $CH_3CSNH_2 + 2H_2O \rightarrow CH_3COO^- + NH_4^+ + H_2S$.

Sections 28.12-28.14. **Exhibits:** Exhibit samples of metallic sulfide. Mention iron pyrites as a source of sulfur but not of iron. Exhibit CS_2. Cover details in Sec. 28.14.

Section 28.15-28.20. **Laboratory Experiment 60,** "Sulfite and Sulfate Ions." Procedures in this experiment may be used for demonstrations of preparation and properties of SO_2 as well as tests for sulfite and sulfate ions. **Demonstrations:** (a) Preparation by methods 1, 3, 4 in Section 28.16. For method 2 refer to Instructor Experiment, Experiment 47. (b) "Wine to water to milk." Pour dil. $KMnO_4$ solution into an "empty" cylinder (small amount of bisulfite of a very small quantity of SO_2 solution) and then into a second cylinder containing "water" [$BaCl_2(aq)$]. Also see Alyea p. 42, 19-17 for another color change. (c) Liquefaction: Pass the dry gas through a U-tube cooled in a Dry Ice-acetone. Under the hood, place a small test tube containing H_2O into a larger tube containing the liquid SO_2. (d) Bleaching: Bleach a flower in SO_2 and restore its color with H_2O_2. (e) Acid anhydride: Test $SO_2(aq)$ with litmus and pH paper. Write ionization equations and the K_i for each. (f) Hydrolysis of slats of sulfurous acid. (g) Stability: Demonstrate that SO_2 neither burns nor supports combustion. Discuss ecology problems of air pollution and smog.

Section 28.21. **Exhibits:** Introduce sulfuric acid by illustrating and/or citing the many uses of H_2SO_4 (Sec. 28.24). Exhibit storage battery, superphosphate fertilizer, rayon, cellophane, etc. **Demonstrations:** *Charring action (dehydration)* 1. Wood splint or cloth dipped into H_2SO_4. 2. Add H_2SO_4 to sugar in a beaker and stir. 2. Removal of oxide coat: Immerse a rusted article in the acid. Identify H_2SO_4 as a "Key Chemical." Quote Aldous Huxley, "Chemists cheer for the flag and H_2SO_4. "Pose the problem of production of sulfuric acid arising from the fact that H_2SO_4 has been used in the production of other acids (both binary and oxyacids). Students suggest that the anhydride be used. Have them decide on the formula of the anhydride. Burn sulfur and then attempt to burn SO_2. This identifies the key problem element, the conversion of SO_2 to SO_3. **Demonstration** Contact process: An evaporating dish covered with a funnel which permits air to be drawn by aspiration over the sulphur which is burning in the dish. The mixture of air and SO_2 is drawn over heated platinized asbestos and the products are drawn through H_2O in a side-arm test tube. The aspirator is attached to the side arm. The H_2O is tested for hydronium and sulfate ions before and after the run. Attention is drawn to the undissolved cloud of SO_3 which did not react with the H_2O. This indicates the need for solution of SO_3 in H_2SO_4. No white cloud forms when, in a second run, a test tube containing H_2SO_4 is substituted for the H_2O. The catalyst may also be V_2O_5 or an electrically heated platinum. See Alyea p. 41 for another version of the contact process and for a lead chamber process demonstration. **References:** See Fowles pp. 216-222 and Experiment 43, "Sulfur to Sulfuric Acid" in *Scientific Experiments in Chemistry.* **Student report:** *Chemistry*, "The Early Days of Sulfuric Acid," 45, Oct. (1972).

Section 28.22. **Activities and demonstrations:** Have a student lift a small bottle of H_2SO_4 and an identical bottle with an identical volume of H_2O. Have student observe oily viscosity as you swirl the acid in a bottle. Carefully heat a sample of the acid with a 360°C thermometer inserted. Demonstrate the correct method of dilution. Have a student feel the outside of the beaker in which this is done or observe an inserted thermometer.

Section 28.26. **Re-demonstrate from Experiment 60.** Contrast test for sulfites and sulfates. **Summary:** Review names, formulas of the three main acids of sulfur, their corresponding salts, and the related hydride and oxides. Have students determine the oxidation state of sulfur in the acids, salts, oxides, and hydride. Review the tests for sulfide, sulfite, and sulfate ions.

Films:

CORF *Sulfur and Its Compounds* (16 mm, 14 min, b-w, c, s)

CORF *Sulfur and Hydrogen Sulfide* (16 mm, 21 min, b-w, c, s)

AIM YF-249 *Sulfur, Its Physical States and Properties* (16 mm, 9 min, c, s)

Sulfur (22 min, s and c), U.S. Bureau of Mines, 4800 Forbes St., Pittsburgh, Pa. Frasch mining operations, uses.

Sulfur Mine at Sea (28 min, s and c), Freeport Sulfur Co. 161 E. 42nd St., N.Y., N.Y. 10017. First offshore mine, use in industry and agriculture.

Transparencies:
HRW Series: CH-7182 Structure of S_8 molecules of sulfur (Fig. 22-2 and Fig. 28-3)
CH-7193 A chain of Sulfur atoms as found in Mu-sulfur (Fig. 28-4)
CH-7194 A flow diagram of the contact process (Fig. 28-9)
K & E 6120 Series
13 Allotropes of Sulfur (3 + 1)
W Slides:
The Oxides of Sulfur
The Properties of Sulfuric Acid
The Preparation and Properties of Hydrogen Sulfide
The Properties of Sulfur
Carbon Disulfide
Filmstrips:
DG 567 Putting Sulfur to Work (40f, c)
DG 695 Chemistry of Group VIA (40 f, c)
General Book References: See Part IV, Bibliography.
Book references for this chapter: Asimov, I., *Building Blocks of the Universe*, Abeland-Schulman, 1961.

TEACHER NOTES CHAPTER 29

Chapter Note: **Activities and Demonstrations:** The chapter topic may be introduced by a variety of approaches such as: (1) show a film: CHEM 4112, *Chemical Families*; CHEM 4160, *A Research Problem: Inert (?) Gas Compounds.* (2) Discuss fluoridation and chlorination of water, TV commercials on dentifrices, pesticides and ecology (DDT, chlorinated biphenyls) or the noble gas compounds of fluorine. The role of theoretical chemistry as an instrument of prediction is typified by the suggestions of Pauling in 1933. A good source is "The Chemistry of the Noble Gases," H. Selig, J. G. Malm, and H. Classen, *Scientific American*, Vol. 210 No. 5, May 1964, pp. 66-77. Student reports may be given based on this or other references. For instance, two reports on Teflon appear in *Chemistry*, Vol. 38, 1965, June and August. (3) Exhibits: Salt, iodized salt, tincture of iodine, Teflon cookware, Clorox, etc. If these activities are not used as an introduction, they may be used at appropriate times during the chapter development.

Experiments that may be used include: Experiment 61, "Preparation and Properties of Chlorine," Experiment 62, "Preparation and Properties of Iodine," Experiment 63, "Halide Ions." Experiments from *Scientific Experiments in Chemistry* are Experiment 58, "Chlorine a Typical Halogen," Experiment 59, "Volumetric Estimation of Chloride in Solution," Experiment 60, "Concentration of Sodium Hypochlorite Solution."

Student reports based on the available booklets from listed commercial sources, i.e., Dow Chemical Co., should be used throughout the chapter.

Section 29.1. **Activities:** Exhibit samples and spatial models. Develop a chalk board and student notebook table like Table 29.1. Give symbol and atomic number. Students supply electron configuration and explain how the atom X becomes the X^- ion leading to principal oxidation number. Mention meaning of halogen. Write electron-dot formula for X_2.

In each instance have students discuss the reasons for variations among the halogens. On the basis of electronegativities and electronegativity differences discuss hydrogen bonding and acid strengths of the hydrogen halides.

Section 29.4. **Exhibit Discussion:** Exhibit fluorspar, fluoride insecticide (Merck Sodium Fluoride includes warning about inhaling, swallowing and antidotes), Freon, spray cans, cryolite, and Teflon-coated cookware. Discuss water fluoridation.

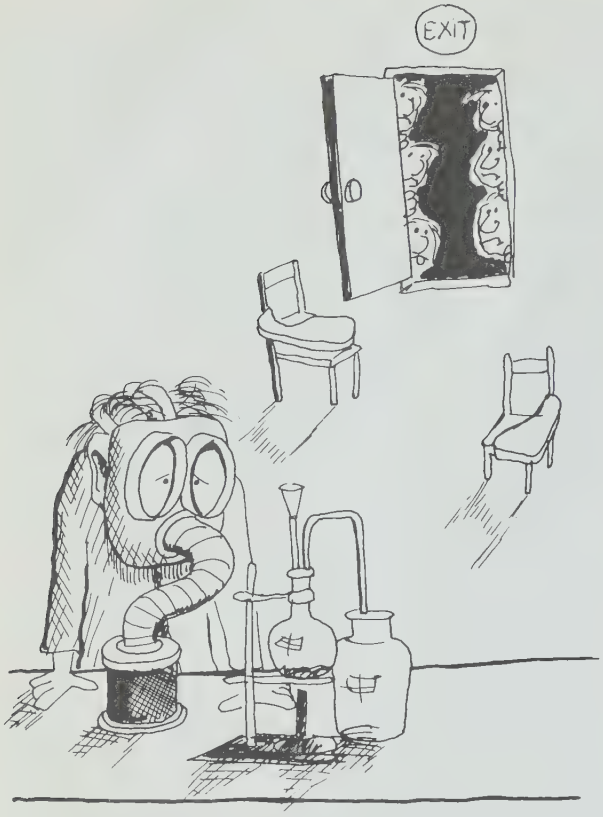
Sections 29.5-29.6. **Activities:** Preparation and etching, see Alyea p. 34, 15-24. Write equations. Display etched glassware and plastic bottle of HF. Discuss hydrogen bonding, ionization, strength, neutralization, and uses of hydrofluoric acid. **References:** *Journal of Chemical Education*, "Strength of Hydrohalogen Acids," 47, 697 (1970), "The Noble Gases and the Periodic Table: Telling it like it was," 46, 569 (1969); *Chemistry*, "Fluorine, a Hostile Element," 42, Feb. (1969); "Teflon" articles 38, June, Aug., (1965); "The Noble Gas Compounds," 37, Jan. (1964).

Sections 29.7-29.8. **Laboratory Experiment 61:** "Preparation and Properties of Chlorine." Exhibit chlorides (Sec. 29.7 and student reference to the Morton booklet, "Salt." **Demonstration:** Demonstrate the electrolysis of brine using a U-tube, and carbon electrodes. Collect hydrogen through outlet tube and test by "bark" test. Use red and blue litmus at both electrodes, test for chlorine by odor and bleaching, test for NaOH with litmus. Develop electrode, net formula, and net ionic equations. Consider or review equations for preparations as oxidation-reduction phenomena and, if desired, balance by redox techniques.

Section 29.9. **Demonstrations:** Use student results in Experiment 61 or re-demonstrate. Refer to method of collection but production of $Cl_2(aq)$ solutions to indicate moderate solubility, and density. Discuss poisonous nature and antidotes for accidental inhalation.

Section 29.11. **Exhibit:** $CHCl_3$, CCl_4, CCl_2: CCl_2 (perchlorethylene), CH_2Cl_2 (paint and varnish removers). Develop representative equations for the demonstrations. Discuss chlorination of drinking water and pools.

Section 29.12. **Demonstrations:** Prepare and collect HCl using apparatus set-up in Fig. 29-5 for Cl_2. Have an extra bottle containing about one inch of H_2O for producing hydrochloric acid. After collecting three



Preparation of chlorine if your laboratory is not equipped with a hood.

bottles of the gas, place the delivery tube about one-quarter inch above the H_2O in the bottle so that students may observe the solution (oily convection currents are observable when the water is viewed from the side). Write equations and discuss other preparations reviewing the burning of hydrogen in chlorine.

Section 29.13. **Demonstrations:** Blow breath across top of bottle of HCl. Test for odor by waft test. Invert a bottle of HCl in H_2O, remove cover plate and note rapid rise of H_2O. Demonstrate the HCl founatin.

Section 29.14. **Demonstrations:** Test a sample of the solution made above with litmus. (The solution is concentrated enough to react with magnesium if not with zinc.) Re-demonstrate protolysis by conductivity and litmus tests of toluene + HCl before and after additions of H_2O. Mix toluene solutions of HCl and NH_3. Bring moistened stoppers from concentrated hydrochloric acid and $NH_3(aq)$ together. Note accumulation of solid on the outside of bottles of hydrochloric acid which may have been stored near $NH_3(aq)$ in stockroom. Wash out a 500-ml cylinder with $NH_3(aq)$. Position the end of the delivery tube from HCl generator at bottom of cylinder and pass in the HCl. A cloud, and solid deposit of NH_4Cl forms on the walls of the cylinder and, on occasion, a perceptible amount of H_2O may be poured from the inverted cylinder. Test HCl(g) with a burning splint for combustibility and for supporting combustion. Write equations for protolysis of HCl (Brønsted acid) with (1) H_2O and (2) NH_3. Write equations for acid-base neutralization.

Section 29.15. **Demonstrations:** Remove coating from brass doorknob or other brass fixture by rubbing with cotton moistened with concentrated hydrochloric acid. Discuss acid concentration in gastric juice and recall or demonstrate the digestion of boiled egg white in pepsin + HCl.

Section 29.17. **Demonstrations:** Demonstrate the test for Cl^- ion. Show the nonspecificity of merely adding $AgNO_3$ by demonstrating precipitation with solutions of carbonate, phosphate, and oxalate. Precipitates are somewhat white, but will dissolve in HNO_3. Develop the chloride ion test sequence by demonstrating solution of AgCl in $NH_3(aq)$ and its reprecipitation with HNO_3. Write the equations. **Laboratory Experiment 63,** "Halide Ions," differentiates AgCl, AgBr, and AgI precipitates. Review Experiment 53 "Separation of Lead (II), Silver, and Mercury (I) Ions," and Experiment 54 "Complex Ions of Copper and Silver." Mention chloride test for developing "latent" fingerprints.

Section 29.19. **Demonstrations and student reports:** Add chlorine water to solution of bromide. Perform shakeout test with perchlorethylene (CS_2, $CHCl_3$, CCl_4). As a control demonstrate shakeout test with bromine water.

Section 29.22. **Student reports:** on Dow booklets: *Bromine and Brominated Compounds* (history and uses). Discuss uses of bromine and bromides (Sec. 29.22): bromine and the gasoline industry—tetraethyl lead, ethylene dibromide, low lead gasoline; bromides as sedatives (exhibit Bromo-Seltzer); AgBr in photography.

Sections 29.24-29.25. **Experiments and Demonstrations:** Laboratory Experiment 62 "Preparation and Properties of Iodine." Use Experiment 62 for demonstrations as desired.

Section 29.26. **Demonstration:** 1. Sublimate a few small iodine crystals in a test tube by heating. Pour out the vapor. Note iodine crystals on the upper part of the test tube. 2. Solubility: Procedure 2, Experiment 62. 3. Starch test for free iodine. 4. Liquefaction of Iodine (Alyea, p. 203). 5. I_2 + white P (*Caution*). 6. Comparative reactivity: antimony + chlorine; antimony heated in deflagrating spoon + $Br_2(g)$; antimony (heated) into hot iodine (g). Replacement of I^- by both chlorine water and bromine water with shakeout test. Refer to Tables 21.3, 21.4 (Standard Oxidation Potentials).

Section 29.28-29.29. **Exhibit:** Tincture of iodine, iodized salt and discuss uses, including uses of radioactive iodine in medicine. **Other Halogen Demonstrations:** *Journal of Chemical Education*, 46, A107-08,

A355-356, A309-310, A451 (1969). Hydrogen halides: Fowles, Experiment 325, 326 for preparation from PBr_3 and PI_3. **Demonstrations:** Place a sample of the sodium or potassium halides in Erlenmeyer flasks. Add H_2SO_4 to each halide and heat. The vapors from the fluoride etch the flask; from the bromide they are red-brown; from the iodide they are violet; and from the chloride, colorless. (HCl will fume if you blow across the top of the flask. The results are representative of the methods for producing HF and HCl and of the smaller resistance to oxidation by the HBr and HI. **Demonstration of interhalide compound** (ICl_3 or $KICl_4$): Place a long strip of filter paper in a 500-ml cylinder. Chlorine enters the cylinder through a glass delivery tube reaching the cylinder bottom. At first, the expected happens (the paper turns brown), but the iodine then reacts with excess chlorine to form the yellow compound(s) from the bottom up (of the paper in the cylinder).

Films:

MLA CHEM 4175 *Bromine—Element from the Sea (16 mm, 22 min,*c,s);
CHEM 4112 *Chemical Families* (16 mm, 22 min, c,s);
CHEM 4124 *Equilibrium* (16 mm, 24 min,c,s);
CHEM 4160 *A Research Problem: Inert (?) Gas Compounds* (16 mm,c,s);
CORF *Preparation and Properties of the Halogens* (16 mm, 32 min,b/w,c,s);
The Halogens (16 mm, 16 min,b/w,c,s);
SUTH *Chlorine—A Representative Halogen* (16 mm, 15 min,c,s);
AIM *Commercial Production of Chlorine* (16 mm, 9 min,c,s);
MH 612000(b/w), 612025(c) *The Family of Halogens* (16 mm, 13 min);
AEC *Chemical Somersault* Free Loan (16 mm, 29 min,b/w,s);
AIM YF-244 *Electronegativity* (16 mm, 4 min,c,s).

Loops:

MLA CHEM Study Film Loops: 168 W 1003(T) or 168 W 1023(K) *Alkali Metal Reactions with Chlorine and Water* (s-8, 4 min,c);
H-E or E E 84-0140/1 *Oxidation and Reduction: The Halogens* (s-8, 4 min,c).

Filmstrips:

DG 632 *The Halogens—Group VII* (40 f, c);
DG 647 *The Noble Gases* (40 f, c).

Transparencies:

HRW CH-7195 Models of some molecules found in Hydrogen Fluoride at room temperature
CH-7117 Periodic Table of Ionization Energies
CH-7125 Periodic Table of Electronegativities
HN2-10528 The Fluorine Molecule
2-10531 The F_2O Molecule
K & E 6120-36 Hydrochloric Acid Reaction (8)
6120-52 Halogens
McW Size and Electronegativity—Group VIIA
Properties of Group VIIA Elements

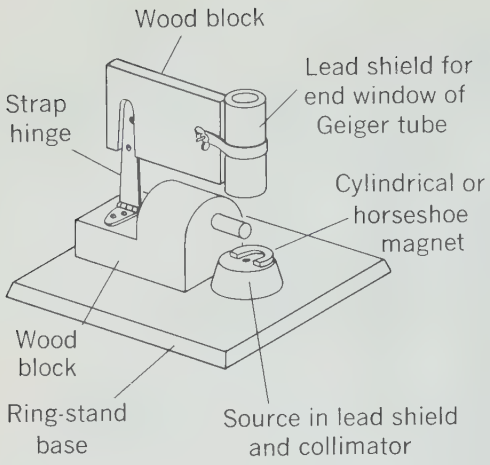
Booklets:

Morton Salt Co., Morton-Norwich Products Inc., 110 North Wacker Drive, Chicago, Ill. 60606
Salt (26 pages) History, Production processes, uses, Research
The Chlorine Institute, 342 Madison Ave., New York, N.Y. 10017
Exceeding All Expectations (36 pages) A short history of Chlorine.
Chlorine Facts (26 pages), Early history, Manufacture today, Safety, Uses (58)

TEACHER NOTES CHAPTER 30

Sections 30.1-30.2. Exhibit: a radiophotograph of an object such as a key, prepared previously, as an example of Becquerel's discovery. Consult *25-year Cumulative Index, Journal of Chemical Education,* (Vol. 1-25) for articles on the Curies.

Sections 30.3-30.4. **Demonstrations:** Refer to Fig. 30-2 and demonstrate the discharge of a Braun electroscope by a radioactive source. Demonstrate the change in count of a Geiger counter exposed to a radioactive source. Test a luminous watch with the Geiger counter. Define half-life. Refer students to Table 30.1 for half-life data. Mention other half-life data: iodine-131, 8 days; phosphorus-34, 14.3 days; strontium-90, 20 years. You may wish to discuss fallout and strontium-90. Demonstrate Wilson cloud chamber (student project or commercial variety, i.e., Cenco 71245). Refer to Table 24.1 for Group II and radium properties. **Half-life demonstration:** Use "Mini-generator" (Welch 2191, 2191A) and directions in laboratory manual, Welch 219 M Experiments in Nucleonics for demonstration(s). Use Table 30.1 for student location of natural radioactive nuclides with atomic number less than 83(Bi). Mention use of rubidium-strontium dating in determining the age of the earth and of moonrocks, and of carbon-14 in dating artifacts. Discuss the generalizations in Sec. 30-3. Ask why astatine and francium are not included among the natural radioactive nuclides. State the use of radon: refer to *Chemistry*, "Latest on Chemistry of Radon," 42, Dec. (1969). **Demonstration:** Separating Beta Rays and Gamma Rays. Provide a collimated beam of radiation by a source in a lead shield. The shield in the illustration was cast in a 120-ml casserole, then a 3/16" central hole was drilled in it for the beam. The magnetic field can be provided by a small horseshoe magnet resting on the flat top of the shield, or by a cylindrical magnet mounted with a pole near the beam or radiation. The detector is an end-window Geiger tube in a lead shield perforated at the bottom. The detector and shield are on a hinged support so that they can be swung about a horizontal axis in the de-



flecting field. **Operation:** In the absence of a magnetic field, the total radiation is counted with the detector directly above the vertical beam. With a magnetic field present, the beta rays are deflected and the count drops to the level of gamma activity alone. The detector may then be swung to find the deflected beam (hand rule), and the beta count made. With luck, the sum of the deflected rays and the undeflected rays should equal the original undeflected beam. The demonstration and illustration are taken from a presentation by George Kapp, James Madison High School, Brooklyn, N.Y. at the October, 1971 Northeast Regional meeting of the National Science Teacher's Association. **Activities:** Start with a chalk board diagram of a single beam of radiation without electric charged plates. Add plates and show three beams which students identify. Supplement Experiment 69 by using beta source (uranyl nitrate) and gamma source (thorium nitrate) and sheets of paper, glass, aluminum and lead to show penetration and shielding effects. Impress students by showing that gamma radiations go through the hand while betas are stopped. A homemade device for magnetic separation of betas is described on text p. 627. **Demonstrations:** Instructor Experiment 5a in Experiment 69 (chemical separation of thorium-234 and proactinium-234 in uranyl nitrate). See the "Mini-generator" Laboratory Manual Welch 219M for related demonstrations.

Sections 30.6–30.7. **Activities and Transparencies:** Illustrate with nuclear equations in stepwise fashion. Define transmutation. Stress the energy concommitant and the effects of alpha and beta emissions on the mass number and atomic number of the product nucleus in each transmutation. Use PP transparency T44, "Radioactive Displacement Rules" which details changes in the number and kinds of nucleons. For beta emission use: neutron → proton + electron. Overlays for PP T44, 44E, and 44-F deal with positron emission (β^+). Write nuclear equations for alpha and beta emissions. For chemical experiments with radioisotopes see Instructor's Experiments 6, 7, Experiment 65.

Section 30.8. **Discuss:** The uses in determining age of earth, moon dust, moon rocks, and the use of carbon-dating. **Student report references:** Garrett, A.B., *The Flash of Genius*, 13, Van Nostrand, 1962. Carbon-14 dating: Willard F. Libby (original article in *Journal of Chemical Education*, 40, 76, 1963); Kamen, M.D., "The Early History of Carbon-14," *Journal of Chemical Education*, 40, 276 (1963).

Sections 30.9–30.10. **Activities:** Define and illustrate mass defect and binding energy. Develop the generalizations pertaining to the stability of a nucleus and nuclides—mass, ratio of neutrons and protons, and the even-odd nature of the number of neutrons and protons. Illustrate as in Sec. 30.9. Discuss the four types of nuclear reactions and the related changes in stability.

Section 30.11. **Student report reference:** Shamos, *Great Experiments in Physics*, Chap. 19, "Ernest Rutherford, Induced Transmutation."

Section 30.12. **Activities:** Write the equation for the Cockcroft-Walton experiment. Calculate the loss in atomic mass units and convert to Mev by multiplying by 931. See Table 5 Appendix, Conversion *Modern Physics*, for interconversion to other mass units such as grams and to other energy units such as calories or kilocalories.

Section 30.14. **Films:** EBF, *Atom Smashers;* EBF *Nuclear Reactors for Research.* **Filmstrip:** DG 631 "Particle Accelerators." **References:** *Scientific American* Offprints: 234 *"The Linear Accelerator,"* 251 *"Particle Accelerators,"* 322 *"The Two-Mile Electron Accelerator."* **Activities:** *Discuss the structure* (Fig. 30-10) and operation of the cyclotron, and the types of particles accelerated. Repeat as desired for the other accelerator in Sec. 30.14. Although the Van de Graff generator is not mentioned, commercial varieties (Welch 1910–1910M) may be used to demonstrate stimulating experiments including some "hair raising" experiments. Student-constructed models may also be available. See *Modern Physics*, pp. 629-634 for discussion of it and other accelerators. The Los Alamos Scientific Laboratory of the University of California, Los Alamos, New Mexico has a series of booklets for student use. Science Study Series, Wilson, R. R., Littaner, R., *Accelerators.* **Transparency:** K & E 6120-86 "Cyclotron."

Sections 30.15–30.16. **Activities:** Refer to the nuclear equation for Chadwick's experiment as neutron production. Discuss fast neutrons, possible nuclear disintegration, production of slow neutrons (deuterium oxide and graphite), capture any unstable nucleus that may result. Develop the equation series for neutron capture by uranium-238. Use the equations in Table 30.2 to illustrate methods for the production of the transuranium elements. **References:** *Chemistry*, Reprint 90 "Plutonium, the Lively Element, Parts I and II"; "Discovery of Element 105," 43, Nov. (1970); "Predicted Properties of Elements 113 and 114," 43,

Nov. (1970); *Scientific American Offprints*: Seaborg, G. T., Perlman, I., "The Synthetic Elements" (I) 242; (II) 243; (III) 293; Seaborg, G. T., *Man-made Transuranium Elements*, Prentice-Hall, 1963; Seaborg, G. T., "Some Recollections of Early Nuclear Chemistry," *Journal of Chemical Education*, 45, 278 (1968).

Section 30.17. Activities: Write equations for the production of cobalt-60 and phosphorus-32. Discuss their uses.

Sections 30.18-30.19. Activities: Write the nuclear equation for fission. Define a chain reaction. Review the equations in Sec. 30.16 for the production of plutonium-239 and mention its use in fission.

Section 30.20. Activities: Use Fig. 30-16 or draw a chalk board equivalent. Discuss details of structure and operation: fuel, moderator, control rods, shield, heat-exchange (coolant).

Section 30.21. Activities: Define fusion and write the nuclear equation for the net change of four hydrogen nuclei to a helium nucleus. Refer to pp. 613-614, *Modern Physics* for details. Identify as a thermonuclear reaction. The Modern Physics reference also provides the equation and energy details for the thermonuclear bomb. Power," 246 "The Stellarator," 340 "The Prospects of Fusion Power." Fowler, T. K. and Post, R. F., "Progress Toward Fusion Power," *Scientific American*, 215, (6) 21 (1966). Freilich, F. G., "Radiation and People." *The Physics Teacher*, 8, 49 (1970). **Questions:** 10-20, 31-40.

Films of Social importance:

DUART SERIES on the Radioisotopes
PMF 5058 *Medical Effects of the Atomic Bomb*
M 443 *Radioactive Waste Disposal*
DOD-DC3-220 *About Fallout*

Scientific American Offprints:

201. Hans A. Bethe WHAT HOLDS THE NUCLEUS TOGETHER September 1953
202. Geoffrey Burbidge & Fred Hoyle: ANTI-MATTER April 1958
210. William A. Fowler THE ORIGIN OF THE ELEMENTS September 1956
213. Murray Gell-Mann & E. P. Rosenbaum ELEMENTARY PARTICLES July 1957
214. Donald A. Glaser THE BUBBLE CHAMBER February 1955
217. Robert Hofstadter THE ATOMIC NUCLEUS July 1956
219. Donald J. Hughes THE NUCLEAR REACTOR AS A RESEARCH INSTRUMENT August 1953
220. P. M. Hurley RADIOACTIVITY AND TIME August 1949
228. Maria G. Mayer THE STRUCTURE OF THE NUCLEUS March 1951
230. Philip Morrison THE NEUTRINO January 1956
232. Philip & Emily Morrison THE NEUTRON October 1951
234. Wolfgang Panofsky THE LINEAR ACCELERATOR October 1954
235. R. E. Peierls MODELS OF THE NUCLEUS January 1959
236. Richard F. Post FUSION POWER December 1957
239. Bruno Rossi WHERE DO COSMIC RAYS COME FROM? September 1953
241. Erwin Schrödinger WHAT IS MATTER? September 1953
242. Glenn T. Seaborg & I. Perlman THE SYNTHETIC ELEMENTS I April 1950
243. Glenn T. Seaborg & Albert Ghiorso THE SYNTHETIC ELEMENTS II December 1956
244. Emilio Segrè & Clyde C. Weigand THE ANTIPROTON June 1956
246. Lyman Spitzer, Jr. THE STELLARATOR October 1958
247. S. B. Treiman THE WEAK INTERACTIONS March 1959
251. Robert R. Wilson PARTICLE ACCELERATORS March 1958
252. Herman Yagoda THE TRACKS OF NUCLEAR PARTICLES May 1956
253. John H. Reynolds THE AGE OF THE ELEMENTS IN THE SOLAR SYSTEM November 1960
261. Victor F. Weisskopf & E. P. Rosenbaum A MODEL OF THE NUCLEUS December 1955
269. Robert E. Marshak THE NUCLEAR FORCE March 1960
282. Gerard K. O'Neill THE SPARK CHAMBER August 1962
290. R. D. Hill RESONANCE PARTICLES January 1963
293. Glenn T. Seaborg & Arnold R. Fritsch THE SYNTHETIC ELEMENTS: III April 1963
296. Geoffrey F. Chew, Murray Gell-Mann & Arthur H. Rosenfeld STRONGLY INTERACTING PARTICLES February 1964
310. Werner H. Wahl & Henry H. Kramer NEUTRON-ACTIVATION ANALYSIS April 1967
311. Hannes Alfvén ANTIMATTER AND COSMOLOGY April 1967
319. Jerome B. Wiesner & Herbert F. York NATIONAL SECURITY AND THE NUCLEAR-TEST BAN October 1964
322. Edward L. Ginzton & William Kirk THE TWO-MILE ELECTRON ACCELERATOR November 1961
323. Gerard K. O'Neill PARTICLE STORAGE RINGS November 1966
324. Leon M. Lederman THE TWO-NEUTRINO EXPERIMENT March 1963
336. Glenn T. Seaborg & Justin L. Bloom FAST BREEDER REACTORS November 1970
340. William C. Gough & Bernard J. Eastlund THE PROSPECTS OF FUSION POWER February 1971
811. Edward S. Deevey, Jr. RADIOCARBON DATING February 1952

Films:

MLA CHEM 4178 *Transuranium Elements* (16 mm, 23 min, c, s)

AEC Audio-Visual Branch, Division of Public Information, U.S. Atomic Energy Commission, Washington, D. C. 20545. Send for listing. Samples of films available are merely listed by name since details for these free loan films (16 mm) are extensive:
Principles of Thermal, Fast and Breeder Reactions (9 min);
Radiation Detection by Ionization (30 min, b/w, s);
Radiation Detection by Scintillation (30 min, b/w, s);
Radiation and the Populations (29 min, b/w, s);
Radiological Safety (30 min, b/w, s);
The Nuclear Witness (28 min, c, s);
Air and Gas Cleaning for Nuclear Energy (30 min, c, s);
Atomic Energy for Space (17 min, c, s);
Atomic Weatherman (18 min, c, s);
Fusion Research (22 min, c, s);
The High Energy People (6 min, c, s);
The Magnetic Bottle (10 min, c, s);
Man and the Atom (59 min, c, s);
Microscope for the Unknown (29 min, b/w, s);
The Worlds Within (29 min, c, s);
Alpha, Beta and Gamma (44 min, b/w, s);
Beta-Ray Spectrometer (7 min, c, s);
Neutron Image Detector (6 min, c, s);

DUART series on The Radioisotope: (Free loan)
PART 1, *Fundamentals of Radioactivity* (16 mm, 59 min, b/w, s);
PART 2, *Properties of Radiation* (16 mm, 68 min, b/w, s);
PART 5, *The Physical Principles of Radiological Safety* (16 mm, 5 min, c, s);
PART 13, *General Sciences* (16 mm, 46 min, b/w, s);

EBF Series on Atomic Energy:
PMF 5058 *Medical Effects of the Atomic Bomb* (16 mm, 32 min, c, s);
DOD-CD_3-220 *About Fallout* (16 mm, 24 min, c, s);
M-443 *Radioactive Waste Disposal* (16 mm, 24 min, c, s);
370 *Atomic Energy* (16 mm, 11 min, b/w, s);
1897 (b/w) 1890 (c), *Atomic Energy—Inside the Atom* (16 mm, 13 min, s);
655 *Atomic Radiation* (16 mm, 13 min, b/w, s);
661 *Carbon Fourteen* (16 mm, 12 min, b/w, s);
Other titles are: *The Atom and Agriculture; The Atom and Industry; Atom Smashers; Atomic Radiation.*
GE *A Is for Atom* (16 mm, 16 min, c, s);
IFB 2, IFB 250 *Conquest of the Atom* (16 mm, 22 min, c, s);
EYE 512 (b/w) 513 (c) *Detectors (Nuclear Radiation)* (16 mm, 15 min, s);
EYE 522 (b/w) 523 (c) *Fallout (Nuclear Radiation)* (16 mm, 15 min, c, s);
EYE 520 (b/w) 521 (c) *Uses in Industry (Nuclear Radiation* (16 mm, 15 min, s);
EYE 518 (b/w) 519 (c) *Uses in Space Studies (Nuclear Radiation)* (16 mm, 17 min, s);
INDU FS-375 *Beyond Uranium* (16 mm, 29 min, b/w, s);
ROB *The Building of the Bomb* (16 mm, 72 min, b/w, s);
IFB 2, IFB 395 *The Discovery of Radioactivity* (16 mm, 15 min, c, s);
MLA YF-227 *Geiger Counter* (16 mm, 7 min, c, s).

Loops:
EBF R80206 (8), S80206 (s-8) *Chain Reaction—Controlled Chain Reaction* (3^- min, c);
EBF R80207 (8), S80207 (s-8) *Critical Size (Nuclear Reactions)* (2^+ min, c);
EAL 80-3346/1 *Radioactivity* (s-8, 4 min, c)
EAL 80-2009/1 *Radioactive Decay* (s-8, 4.20 min, c);
EAL 80-2017/1 *Scintillation Spectrometry* (s-8, 3.10 min, c);
ICF 59645 *Biological Effects of Nuclear Radiation* (s-8, 4 min, c).

Filmstrips:
DG 568 *Using Atomic Energy for Electric Power* (40 f, c)
DG 582 *What's in the Atom?* (40 f, c)
DG 593 *What Is a Neutron?* (40 f, c)
DG 595 *Atomic Detectives* (40 f, c)
DG 631 *Particle Accelerators* (40 f, c)
DG 641 *Radiation Detection* (40 f, c)

Transparencies:
HRW CH-7196 The effects of a magnet on different types of radiation
HRW CH-7197 The Uranium decay series
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